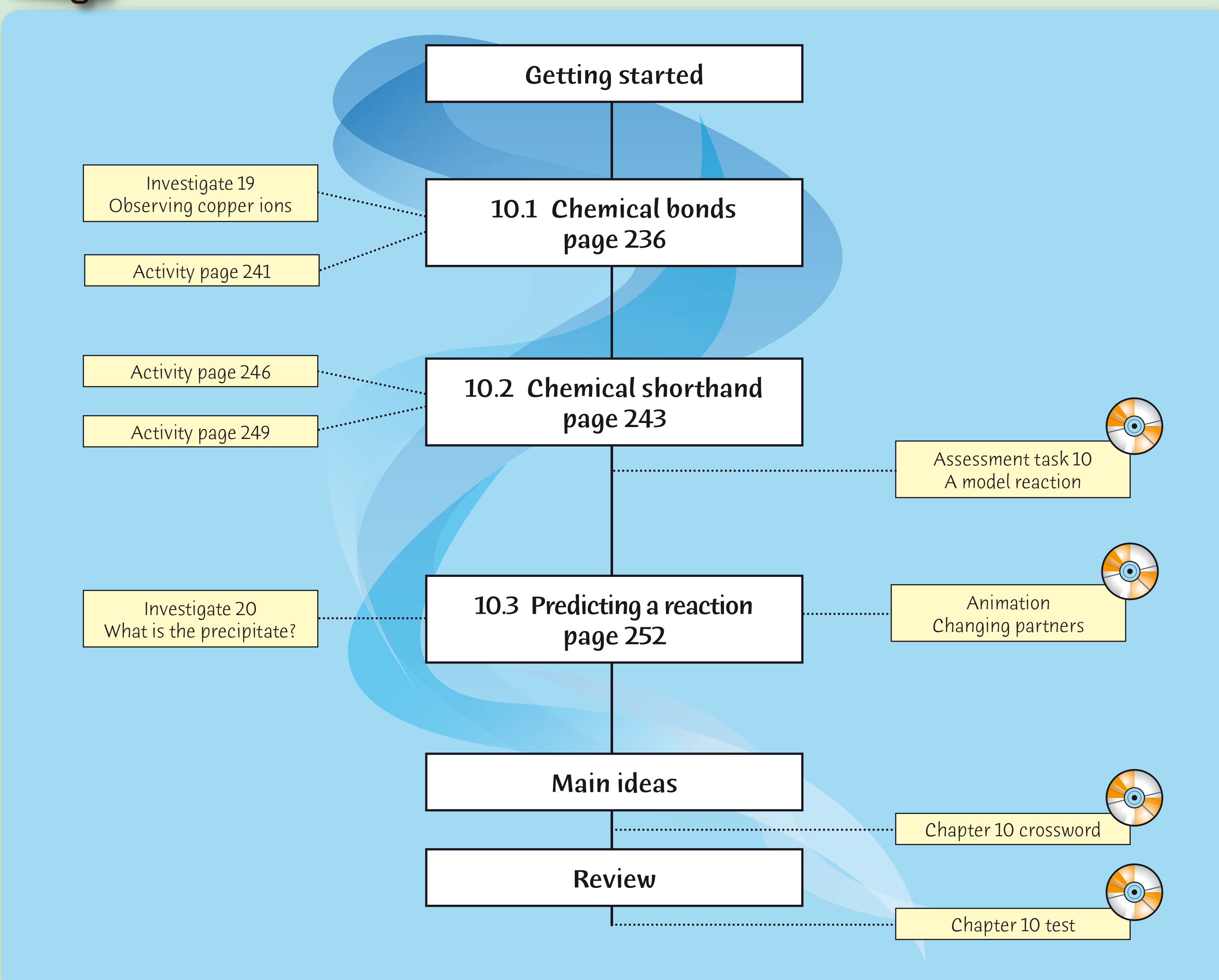
10 Explaining reactions

Planning page



Essential Learnings for Chapter 10

Essential Learnings	*References*		
	Student book (page number)	***Workbook (page number)***	***Teacher Edition CD (Assessment task)***
Knowledge and understanding *Natural and processed materials* Matter can be classified according to its structure	pp. 236–256	pp. 78–84	Assessment task 10 A model reaction
Chemical reactions can be described using word and balanced equations	pp. 243–251	pp. 83–84	Assessment task 10 A model reaction
Ways of working Conduct and apply safety audits and identify and manage risks	Investigate 19 p. 239 Investigate 20 pp. 253–255		
Identify problems and issues, formulate scientific questions and design investigations		pp. 77–78 p. 85	

QSA Science Essential Learnings by the end of Year 9

Vocabulary

ammonium hydroxide
aqueous
copper sulfate
covalent
crystal lattice
diatomic
hydrochloric acid
ionic equation
ions
lead nitrate
magnesium oxide
metallic bond
potassium iodide
precipitate
valency
zinc chloride

Focus for learning

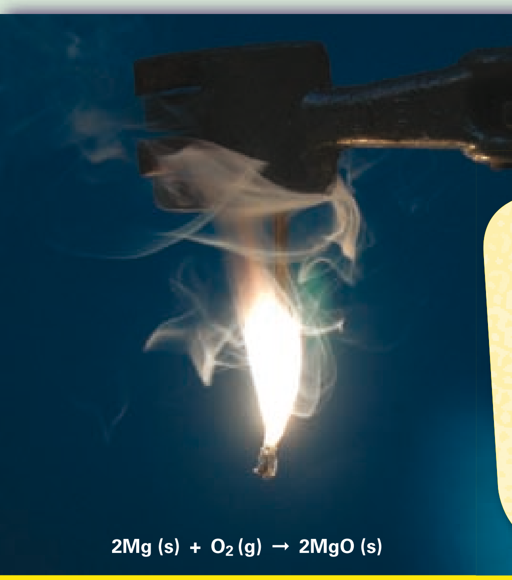
Suggest inferences to explain what happens when you mix lead nitrate and potassium iodide (page 235).

Equipment and chemicals (per group)

Investigate 19 page 239	clean copper strip (approx 4 cm × 1 cm), clean strip of copper, brass or bronze mesh (approx 2 cm × 1 cm), 10 mL of ammonium sulfate solution, petri dish, 2 connecting wires with alligator clips, power pack
Activity page 241	plastic electrostatics rod
Activity page 246	photocopies of ion models (on paper or light cardboard), scissors, glue
Activity page 249	photocopies of molecule models (on paper or light cardboard), scissors, coloured pencils or pens, glue, molecular models kit (optional)
Investigate 20 pages 253–255	**Part A:** 2 test tubes, test tube rack, lead nitrate solution (0.1 M), potassium iodide solution (0.1 M), piece of filter paper, filter funnel, stand and ring clamp, small beaker, wash bottle containing water, watch glass, disposable gloves (optional) **Part B:** lead iodide from Part A, 2 small test tubes, test tube holder, spatula (narrow type), burner and heatproof mat, matches, cotton wool, white tile

Special preparations

Investigate 19 page 239	To make ammonium sulfate solution, dissolve 66 g of $(NH_4)_2SO_4$ in 450 mL water and add 50 mL concentrated ammonia.



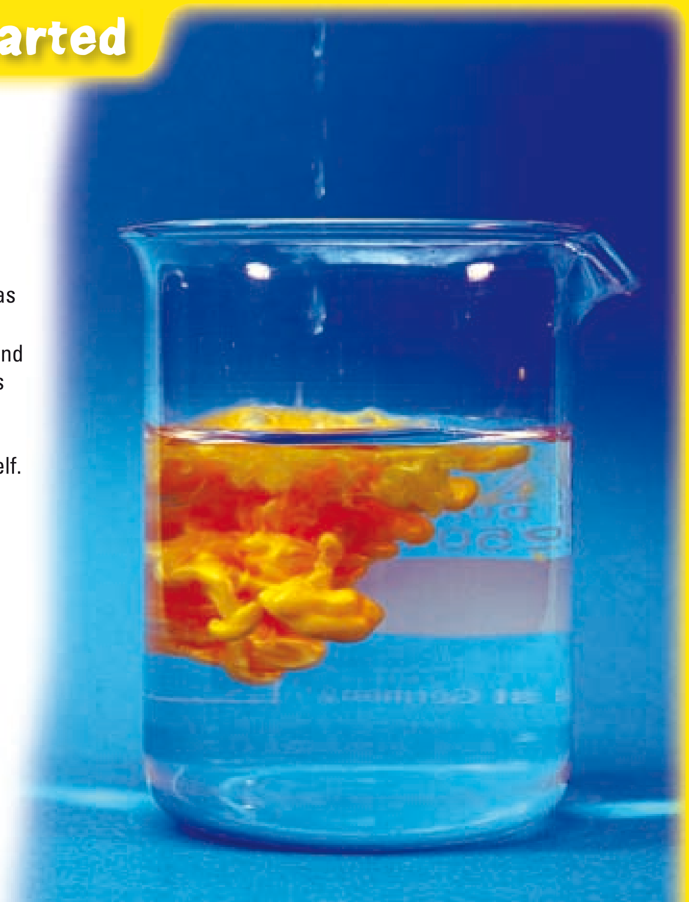
10 Explaining reactions

Getting Started

The photo shows what happened when the teacher mixed two colourless liquids.

- Describe what happened.
- Has there been a chemical reaction? How do you know?
- How could you explain what has happened in terms of atoms?
- Why is it important to understand what happens when chemicals are mixed?

In this chapter you will be able to investigate this reaction for yourself. You will also learn how to explain chemical reactions in terms of atoms, molecules and ions.



Starting point

1 Students could work in pairs to examine the Getting Started questions before discussing them together with the class. What can they remember from their previous studies in chemistry? Consider doing a simple demonstration where calcium metal is added to water. Add a tiny handful or lump of the calcium metal to a large beaker of water that contains a few drops of phenolphthalein. Discuss why phenolphthalein is added. The water becomes cloudy due to the formation of limewater (calcium hydroxide solution) and the pink colour indicates the solution is basic. Hydrogen gas is produced during the reaction—a 'sizzling' noise can be heard—so make sure to use a well-ventilated area to allow the gas to dissipate. The beaker becomes hot because of the release of energy (exothermic) during the reaction. The reaction is:

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g).$$

2 After the demonstration, students can write down what happened in the context of chemistry. It might be helpful to give them a word list to use. Suggested words are: *atoms*, *molecules*, *elements*, *solid*, *liquid*, *gas*, *reactants*, *products*, *chemical reaction*, *chemical formula*, *exothermic*, *solution*, *basic*. Choose students to read aloud their explanation, and from this determine how much revision is necessary. (The demonstration can be performed again in Chapter 11, to show the reactive properties of alkaline earth metals (Group II). For further details, see page 268.)

3 Consider giving the class a pre-test to establish known facts and identify areas which need to be covered in greater detail. A good pre-test should be carefully planned. Test key points that you think the students should already know and that you would like them to understand by the end of the chapter. You may like to give the same test again at the end of the chapter so that students can evaluate their own learning progress. A simple multiple-choice or fill-in-the-gaps test is quick and effective.

4 Revise the difference between a substance changing physically and chemically. See *ScienceWorld 1* pages 44–45 for more details.

10.1 Chemical bonds

Ions

Why is it that salt solution conducts electricity yet sugar solution does not? You can explain this in terms of the particles they contain. Salt solution contains sodium **ions** and chloride ions, which can carry an electric current. On the other hand, sugar solution contains uncharged sugar molecules which cannot carry an electric current.

As you know from previous studies, atoms consist of a nucleus containing positive protons and neutral neutrons, surrounded by negative electrons. The number of electrons is the same as the number of protons, so the atom has no charge. In chemical reactions it is only the outermost electrons that are involved.

Atoms of metals tend to lose electrons. You will find out the reason for this in the next chapter. For example, a sodium atom can lose one electron to form an ion with a single positive charge Na^+, as shown below. Copper atoms lose two electrons to form Cu^{2+} ions with two positive charges, and aluminium loses three electrons to form Al^{3+}.

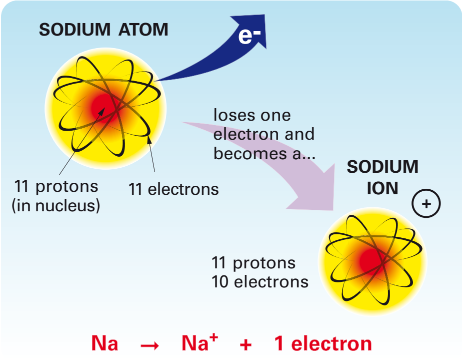


Fig 2 How sodium ions are formed

In contrast to metals, atoms of non-metals tend to form ions by gaining extra electrons. For example, chlorine atoms form negative *chloride* ions Cl^-.

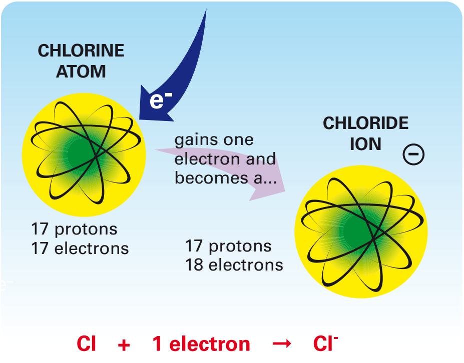


Fig 3 How chloride ions are formed

We can summarise what we have learnt about ions as follows.

METAL ATOMS LOSE ELECTRONS → TO BECOME → POSITIVE IONS

For example:

sodium ($Na \rightarrow Na^+$ + 1 electron)

copper ($Cu \rightarrow Cu^{2+}$ + 2 electrons)

aluminium ($Al \rightarrow Al^{3+}$ + 3 electrons)

Note: Hydrogen behaves like a metal and loses an electron to form H^+ ions.

NON-METAL ATOMS GAIN ELECTRONS → TO BECOME → NEGATIVE IONS

For example:

chlorine (Cl + 1 electron $\rightarrow Cl^-$)

bromine (Br + 1 electron $\rightarrow Br^-$)

iodine (I + 1 electron $\rightarrow I^-$)

Atoms are not usually found on their own. Rather, atoms join together to form molecules such as water H_2O. The atoms in these molecules are held together by the attractive forces of **chemical bonds**.

Hints and tips

You might like to briefly explain to students why ionic bonding occurs.

- One atom transfers electrons to another atom so that both atoms attain a complete outer electron shell (energy level). Individual atoms are unstable unless their outer electron shell (level) is filled. The electrons in this shell are called valence electrons.
- During the formation of an ionic bond, one atom gains electrons and the other atom loses electrons, thus becoming ions.
- The two ions have opposite charges and attract each other. This force of attraction is strong enough to hold the ions together, forming an ionic bond.

This is explained in greater detail in Chapter 11.

Learning experience

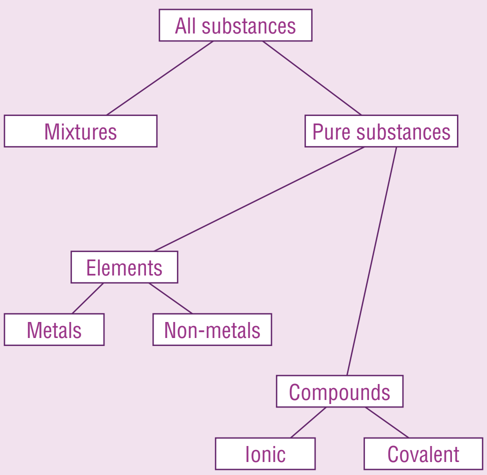
Students could compile their own glossary of terms as they progress through the chapter. They need to write the definition of each term in their own words and give an example of what it is, or how the idea is applied. After each new term has been added, students could check the glossary at the end of the textbook. However, make sure they avoid copying it. The glossary could be in the form of diagrams, or as an audio file that can be played on an MP3 or MP4 player. ESL students may find the audio file helpful so they can hear each word pronounced correctly.

Learning experience

Colourful posters and flow charts are always a good way to summarise and simplify information. Get the class to make posters to display around the room using the information in section 10.1. Have them add some interesting facts about ionic and covalent bonding. Allow students to do further research about each type of bonding. Encyclopaedias, chemistry books or the internet are good resources.

Learning experience

Develop a classification chart for all substances (matter) with the class. Give students time throughout the next few lessons to add to it. (The chart could be continued into Chapter 11.) Suggest they add simple explanations, diagrams or examples to the chart. This is a good activity to do using an interactive whiteboard as you can save it for later use and then print copies of the final chart to give to the students. Alternatively, a poster could be made displaying the information and hung in the room. The following diagram might be a starting point.

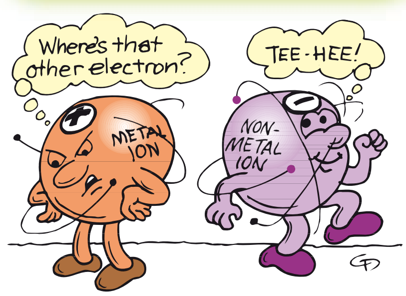


There are two main types of chemical bonds—ionic and covalent. Both kinds of bonds result from the behaviour of the outermost electrons in atoms. Some of these electrons may be lost, some may be gained, or they may be shared. It depends on the kinds of atoms.

Ionic bonds

Metals tend to lose electrons, and non-metals tend to gain electrons. So when a metal reacts with a non-metal, electrons are transferred from the metal to the non-metal, forming positive metal ions and negative non-metal ions.

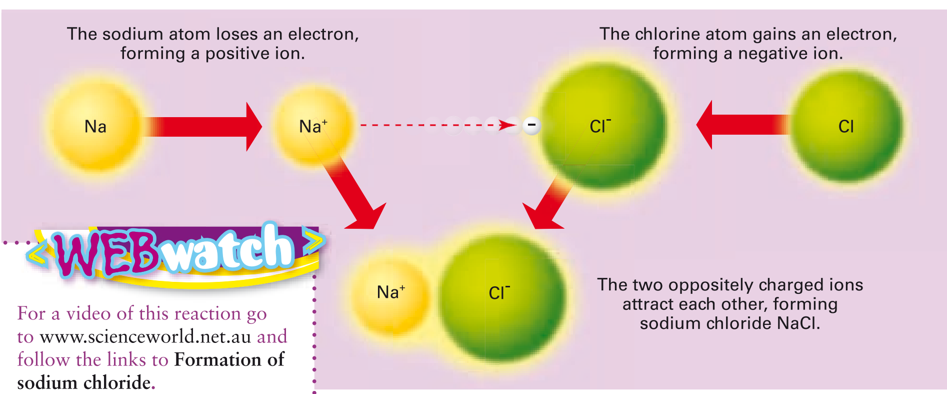
IONIC BOND Ions formed



Ordinary table salt has the chemical name sodium chloride and has the formula NaCl. It is made from the metal sodium and the non-metal chlorine. When sodium and chlorine react, each sodium atom loses one electron to form a positive sodium ion Na^+, and each chlorine atom gains an electron to form a negative chloride ion Cl^-. Opposite charges always attract each other, so the positive sodium ions and the negative chloride ions attract each other. This mutual attraction holds the ions together in what is called an **ionic bond**. See the diagram below.

A salt crystal contains millions and millions of sodium ions and chloride ions, and each ion attracts the oppositely charged ions around it. It is not possible to identify a molecule of sodium chloride because every ion belongs to every other ion surrounding it. There is a regular pattern of positive and negative ions called a *crystal lattice*, as shown in Fig 6 on the next page. If your school has a model of the sodium chloride lattice, notice how the sodium and chloride ions are arranged.

Sodium chloride is called an *ionic compound*. This is because it is made of ions held together by ionic bonds. In solid form, it will not conduct an electric current because the ions are locked into the crystal lattice and cannot move. However, when you dissolve sodium chloride in water, the sodium ions and the chloride ions break apart and spread throughout the water. These ions are free to move and conduct an electric current. The ions carry the electric current through the solution. This is why a salt solution conducts electricity. All ionic compounds conduct electricity when dissolved in water. See Fig 7 on the next page.



WEBwatch

For a video of this reaction go to www.scienceworld.net.au and follow the links to **Formation of sodium chloride.**

Hints and tips

- Small groups of atoms chemically bonded together are called molecules. Water, sugar and carbon dioxide are examples of molecules.
- An ionic lattice (crystal lattice) is when large ionic structures are held together in a regular pattern (array). In most ionic compounds the negative ions (anions) are much larger than the positive ions (cations). These oppositely charged ions are arranged in a regular array forming ionic lattices. As a result, ionic compounds often form crystals—grains of sodium chloride (table salt) are an example of this.

Learning experience

Do the Webwatch if you have access to computers in class, or it can be given as a homework exercise. See if students can use the internet to find some other animations relating to ionic bonding. Have them compile an electronic list of web addresses and their own rating of the quality of each site. Having the list in electronic form means you will have easy access to viewing each animation with the class if access to the internet is available. The same exercise can be done later for metallic and covalent bonding.

Learning experience

Show students how ionic compounds do not conduct electricity in solid form, but do in solution. Place some sodium chloride crystals into a dry beaker, connect a 2.5 V light bulb to a power pack, then put the ends of the leads into the beaker. The leads should not touch each other, but should be partially immersed in the crystals. When you turn the power pack on, the globe will not glow. Repeat the experiment but this time for sugar. Again, the globe will not glow. Make up some sodium chloride solution and place the electrodes into the solution. The globe will glow. Do the same for sugar solution. The globe will not glow.

Hints and tips

A cation is a positive ion and an anion is a negative ion. Although ions are charged particles, ionic compounds can only conduct electricity if their ions are free to move. In solid form they are not free to move, so solid ionic compounds do not conduct electricity. However, they do conduct electricity when they are dissolved in water or melted, allowing their ions to move.

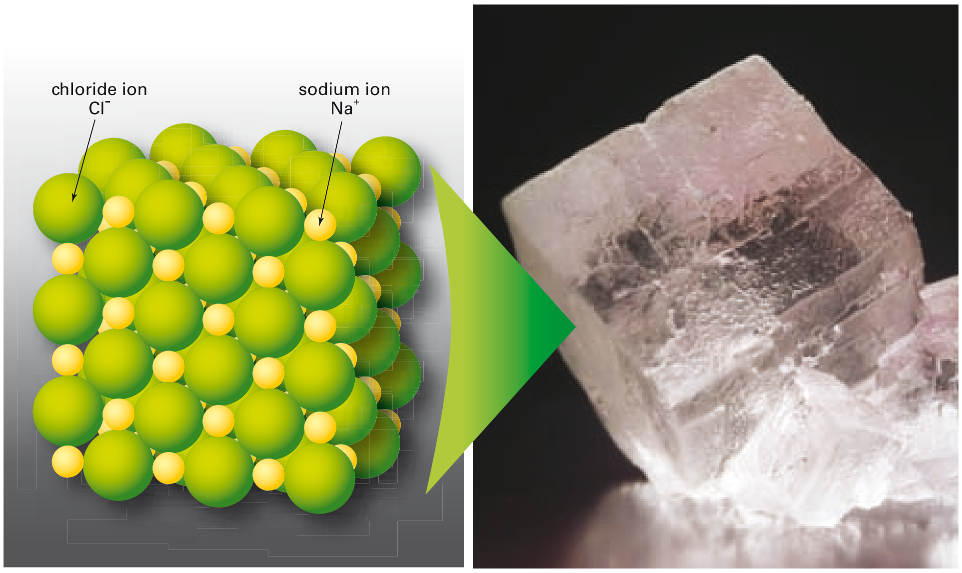


Fig 6 Sodium chloride crystals and their lattice of sodium chloride ions. The positive and negative charges balance each other, so there is no overall charge.

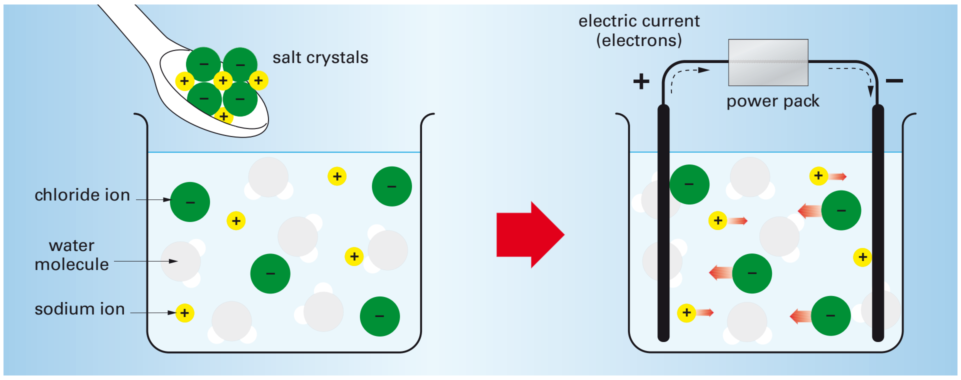


Fig 7 Sodium and chloride ions break apart in the solution. The ions carry the electric current through the solution, and electrons carry it through the wires to and from the power pack.

Individual atoms and ions are of course too small to see. However, you can observe them when there are millions and millions of them together, say in a piece of copper metal. When these copper atoms lose electrons to become copper ions, you can see the results, as in Investigate 19 on the next page.

Learning experience

To make the lesson more interesting, have students make a giant model of ionic lattices. You could start with a sodium chloride lattice, using two sizes of polystyrene balls, two different paint colours, and skewers. Organise the students so that enough ions of each type are made. Appoint one student to direct other class members on how to join the ions together so that a lattice is formed.

Learning experience

Students could write a creative story about their journey into the depths of a grain of salt (sodium chloride crystal). They need to imagine they have shrunk to a size smaller than a sodium ion. Have them explore the world of salt by describing what they see (structure, particle sizes, etc), its properties, how it behaves as a solid, what happens when it is dissolved in water, and any other notable features. You could give them the start of a story and they could finish it. Remind students of the importance of being scientifically accurate as well as creative. Alternatively, you could ask them to write a song or poem describing a sodium chloride crystal.

Investigate

19 OBSERVING COPPER IONS

Aim

To observe the formation of copper ions in a solution.

Materials

- clean copper strip, approximately 4 cm × 1 cm
- clean strip of copper, brass or bronze mesh, approx 2 cm × 1 cm
- 10 mL of ammonium nitrate solution (40 g of NH_4NO_3 in 450 mL water and 50 mL concentrated ammonia)
- petri dish
- 2 connecting wires, with alligator clips
- power pack

Planning and Safety Check

- Your teacher may set up this investigation on an overhead projector so that you can see what happens on a large screen.
- To explain your observations you need to answer the Discussion.

Method

1 Set up the copper strip and copper mesh in a petri dish on a sheet of white paper, as shown. The copper mesh should be about 5 cm from the copper strip. Use clean alligator clips to attach the strip and mesh to the side of the dish.

2 Connect the set-up to the power pack, making sure the copper mesh is connected to the positive (+) DC terminal of the power pack.

3 Half fill the petri dish with the ammonium sulfate solution. Then turn on the power pack to 6 volts DC.

Record your observations of what happens in the next few minutes.

Discussion

1 Towards which terminal (positive or negative) did the blue colour move?

2 Suppose the blue colour is due to copper ions. Which charge would these ions have—positive or negative? Why?

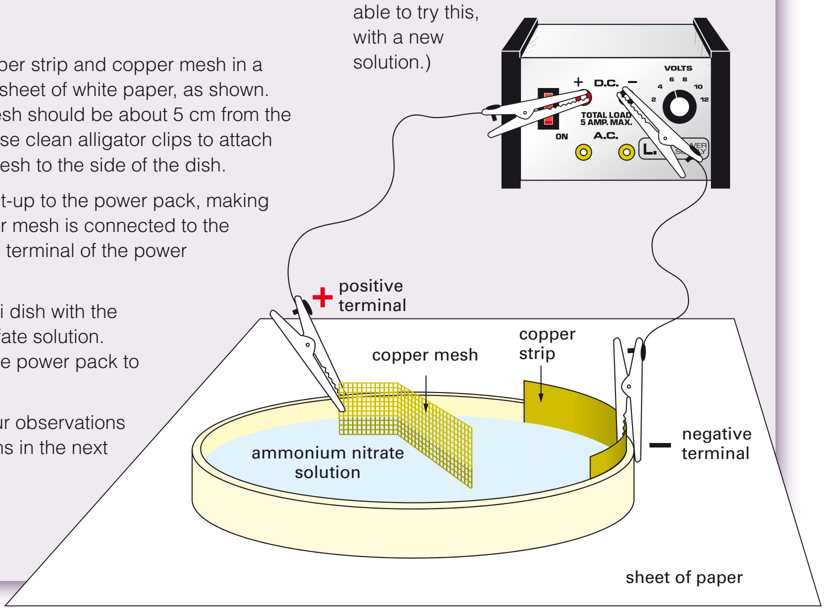
3 Infer where the copper ions came from.

4 In forming ions, did the copper atoms lose electrons or gain electrons? Write an equation for this (see page 236).

5 Although some blue colour was formed on the other side of the copper mesh, it didn't form streaks. Try to explain this.

6 Suppose you added a second copper strip on the other side of the mesh, and then connected both it and the other strip to the negative terminal of the power pack. What do you predict would happen? Explain.

7 Predict what would happen if you reversed the connections to the power pack. (You may be able to try this, with a new solution.)



Lab notes

- The use of ammonium nitrate is now banned in schools in Australia. Ammonium sulfate can be used as a replacement.
- Check that the terminals and clips are clean to ensure a good electrical connection.
- Use a bench mat under the dish in case of spillage.
- This works well as a demonstration using an overhead projector in a darkened room.
- If you are experiencing difficulty with obtaining copper mesh, copper strips with nail-punched holes in it can be used.
- The dark blue colour is due to the formation of a tetraammine copper complex: $Cu(NH_3)_4^{2+}(aq)$

Hints and tips

- Remind students that electric current is the flow of charge. In metals, electrons can move freely, making them good conductors of electricity and heat. Conduction of heat occurs by the vibrations of the positive ions as well as by the freely moving electrons. These freely moving electrons are able to transfer energy (heat) much faster than just vibrations in bonds, as in non-metal solids.
- Consider briefly explaining the periodic table and point out which elements are metals and which are non-metals. An example of the periodic table can be found on pages 260–261. Visual learners may be better able to grasp the concepts of ionic and covalent bonding if valencies have been discussed (see page 243). Reinforce that covalent bonds form between non-metal atoms. Also, be sure to check that students understand the difference between polar and non-polar covalent bonding. Put a list of covalent molecules on the board and have students predict which are polar and which are non-polar.

Learning experience

Students could make a chart comparing and contrasting ionic, metallic and covalent bonding. For each type of bonding, have them write an explanation with labelled diagrams, and give some examples. Their explanations need to be simple enough that a Year 8 student could understand them.

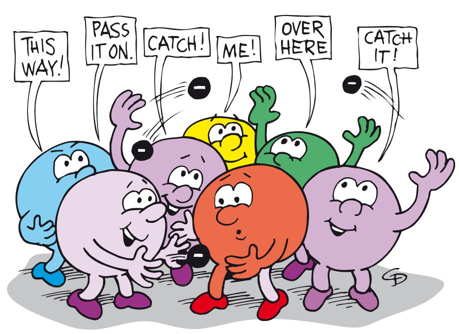
extra for experts

Metallic bonds

Pure metals such as copper, gold and iron are elements and contain only one type of atom. The outermost electrons around these atoms are very weakly held, and they are easily lost. As a result, inside any piece of metal there are a lot of positively charged metal ions that have lost one or more electrons. These positive metal ions are surrounded by a 'sea' of electrons that belong to no particular atom and are juggled about rapidly.

The positively charged metal ions and the negatively charged electrons attract each other, and this mutual attraction is called a *metallic bond*. It is because of these freely moving electrons that metals are such good conductors of electricity and heat.

METALLIC BOND
Free electrons



Questions

1 Copper atoms have two weakly held electrons. Draw a two-dimensional diagram of a tiny piece of copper that contains ten copper atoms. How many free electrons are there?

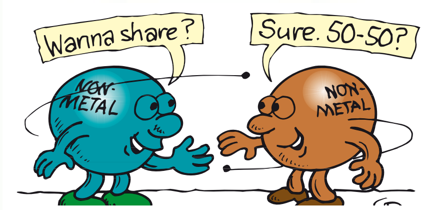
2 Use your diagram to explain why metals are such good conductors of electricity.

Covalent bonds

Non-metal atoms tend to gain electrons. So when two non-metal atoms come together they will both tend to gain electrons, without either of them losing electrons. They can do this by *sharing* electrons to form a **covalent bond**.

When two hydrogen atoms form a molecule of hydrogen H_2, one electron from each atom is shared by the other. Similarly, two chlorine atoms share an electron pair to form a molecule of chlorine gas Cl_2. So, a covalent bond is a shared electron pair holding two atoms together.

COVALENT BOND
Electrons shared



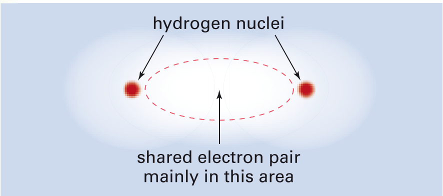


Fig 10 In a hydrogen molecule the hydrogen atoms share electrons equally.

Some non-metal atoms can form more than one bond. For example, oxygen can form bonds with two hydrogen atoms to form water H_2O. Nitrogen can form bonds with three hydrogen atoms to form ammonia NH_3. And carbon can form bonds with four hydrogen atoms to form methane CH_4. It can also form bonds with other carbon atoms (see Fig 12 on page 241). Compounds containing covalent bonds are called *covalent compounds*.

Learning experience

Students could do a similar exercise to the one outlined in the Learning experience on page 238, but make a giant covalent lattice or molecule, such as silica (SiO_2) or sucrose ($C_{12}H_{22}O_{11}$). A variety of different-sized polystyrene balls and coloured paints are useful materials. Alternatively, pairs of students could make some simple covalent molecular models to be displayed around the room. You will probably need to provide them with a list of covalent compounds to choose from.

Homework

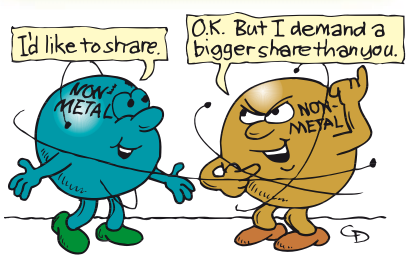
- Prepare some half-page worksheets for the students in which they have to count how many atoms there are of each element in various molecules. Make some of them tricky to challenge students.
- Construct the sheets as a table like the one shown below. Other possible substances for the table are ammonia (NH_3), ammonium chloride (NH_4Cl), carbon dioxide (CO_2), copper (Cu), copper chloride ($CuCl_2$), hydrochloric acid (HCl), methane (CH_4), oxygen (O_2), sodium chloride (NaCl), sodium hydrogen carbonate ($NaHCO_3$), sulfur dioxide (SO_2), sucrose ($C_{12}H_{22}O_{11}$) and water (H_2O).
- Prepare at least two different sheets and rotate them among the students. The sheets can be corrected in class by getting students to swap them and check. If any of their answers differ, they are to confer with each other.

Substance	Formula	Names of elements	No. of atoms of each element	Total no. of atoms in one molecule	Type of bonding
aluminium bromide	$AlBr_3$	aluminium bromine	1 3	4	ionic

Sugar is a covalent compound containing carbon, hydrogen and oxygen. It consists of large molecules with the formula $C_{12}H_{22}O_{11}$. When you dissolve sugar in water, these molecules separate and spread throughout the solution. However, because the molecules do not form ions, the solution does not conduct an electric current. Similarly all covalent compounds are non-conductors of electricity.

Sometimes non-metals don't share their electrons equally. For example, in the covalent bonds between hydrogen and oxygen in water molecules, the oxygen atom tends to get more than its share of electrons because it is 'greedier' for electrons than hydrogen is. As a result the oxygen atom has a slight negative charge and the hydrogen atoms have a slight positive charge. So the molecule has one positive end or pole, and two negative ends or poles. For this reason the bond is called a *polar* covalent bond.

POLAR COVALENT BOND
Unequal sharing of electrons



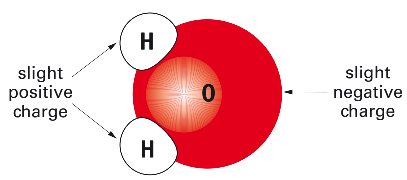


Fig 11 Water molecules are polar because the bonding electrons are pulled slightly towards the oxygen atom. The positive and negative charges balance each other, so the molecule has no overall charge.

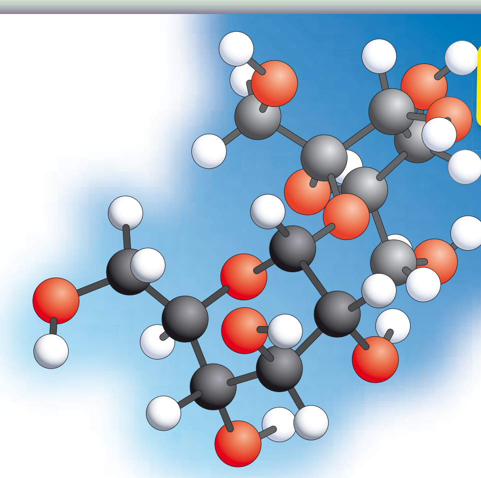
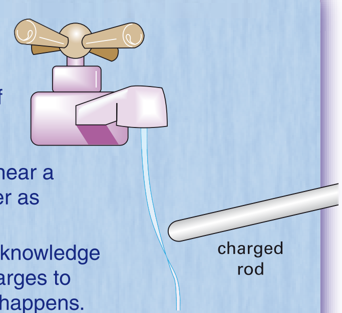


Fig 12 Sucrose (cane sugar) is a covalent compound with the formula $C_{12}H_{22}O_{11}$. In this model carbon atoms are black, hydrogen white and oxygen red.

Bonds between different non-metals are always polar, since no two non-metals are equally greedy for electrons. On the other hand, bonds between atoms of the same non-metal are non-polar: for example the two oxygen atoms in an oxygen molecule. In this case there is equal sharing of the electrons.

Activity

You can demonstrate the polarity of water by bringing a charged rod near a trickle of water as shown.



Use your knowledge of electric charges to explain what happens.

WEBwatch

For a summary of the different types of bonds, go to www.scienceworld.net.au and follow the links to **Dog bone bonds.**

Hints and tips

Give the class a quick quiz based on the material they have already learned. This will indicate any concepts that need to be revised. Ask the students to write the answers only (no need for the questions).

Activity notes

Focus on 'how we know' rather than 'what we know', as this enables students to constructively develop their own knowledge. For example, focus on a question such as 'How do we know water molecules are polar?', rather than 'Water molecules are polar'. Get the students to ask questions, come up with ideas and gather evidence to support their ideas, rather than simply listing facts. Let them try this style of questioning with this activity.

Homework

Students could do a similar activity to the one outlined in the Learning experience on page 237, where they surf the internet to find some animations, but this time relating to covalent bonding. Have them compile an electronic list of web addresses and provide their own rating of the quality and usefulness of the animations. Any animations that have been highly recommended by students could be viewed by the class.

Learning experience

A great activity is to get students to make a cartoon strip to explain one type of bonding. Encourage students to use bright colours and bold images. The cartoons could be collated to form a class comic book or scanned into a computer and presented electronically. Alternatively, students could write poems describing each type of bonding.

Learning experience

Do the Webwatch if you have access to computers in class, or give it as a homework exercise.

Check! solutions

1 a An ion is a charged *atom*.
  b Metal atoms form *positive* ions.
  c Non-metal atoms form *negative* ions.
  d Compounds formed from positive and negative ions are called *ionic* compounds.
  e Covalent compounds form when a *non-metal* atom combines with a non-metal atom.
  f Covalent bonds occur when electrons are *shared* between atoms.

2 a A magnesium atom has 12 electrons.
  b A magnesium ion has 10 electrons.
  c There are 12 protons in the nucleus of a magnesium ion.
  d An electron has a single negative charge.
  e A proton has a single positive charge.
  f A magnesium ion has a charge of 2+.
  g Each magnesium ion has this charge because it has lost two electrons.

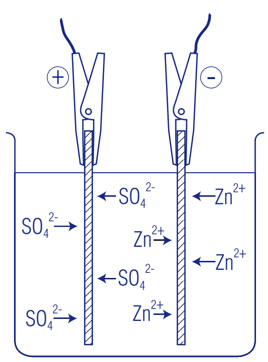
3 Each silver atom loses one electron to become a silver ion.

4 The symbol for an aluminium ion is Al^{3+}.

5 Each bromine atom tends to gain an electron to become an ion with a single negative charge. You know this because it is a non-metal, similar to chlorine.

6 a You would expect this compound (MgO) to be ionic because it contains a metal and a non-metal.
  b The magnesium ions would be positive and the oxygen ions would be negative.

7 a



The zinc ions are positive and will be attracted to the negative electrode. The sulfate ions are negative and will be attracted to the positive electrode.

  b If the connections are reversed then the ions will move in the opposite directions.

Check!

1 Copy and complete the following sentences
  a An ion is a charged ______.
  b Metal atoms form ______ ions.
  c Non-metal atoms form ______ ions.
  d Compounds formed from positive and negative ions are called ______ compounds.
  e Covalent compounds form when a non-metal atom combines with a ______ atom.
  f Covalent bonds occur when electrons are ______ between atoms.

2 A magnesium atom has 12 protons in its nucleus. It loses two electrons to form a magnesium ion.
  a How many electrons does a magnesium atom have?
  b How many electrons does a magnesium ion have?
  c How many protons are there in the nucleus of a magnesium ion?
  d What charge does an electron have?
  e What charge does a proton have?
  f What charge does a magnesium ion have?
  g Explain why a magnesium ion has this charge.

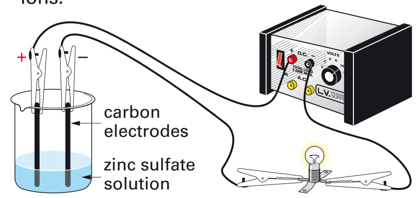
3 The symbol for a silver ion is Ag^+. How many electrons does a silver atom lose to become a silver ion?

4 An aluminium atom Al can lose three electrons. What is the symbol for an aluminium ion?

5 Bromine is a non-metal similar to chlorine. Predict whether bromine atoms tend to lose electrons or gain them. Explain your answer.

6 Magnesium burns in air to form the compound magnesium oxide.
  a Would you expect this compound to be ionic or covalent?
  b Which atoms would be positive and which would be negative?

7 Electricity is passed through a zinc sulfate solution containing zinc ions and sulfate ions.



  a Draw a simple diagram showing what you predict will happen to the zinc and sulfate ions.
  b Predict what will happen if the connections to the power pack are reversed.

challenge

1 What would need to happen for an oxide ion O^{2-} to become an oxygen atom?

2 Copy and complete this summary of chemical bonding.

Type of bond	Kinds of atoms	Electrons are ...	Result
	non-metals		molecules
		lost and gained	

3 Explain why ionic compounds, which consist of electrically charged ions packed together, are electrically neutral. Use sodium chloride as an example.

4 Copper ions are never found on their own. Why is this?

5 Which of the following molecules would you expect to be polar: ammonia (NH_3), chlorine (Cl_2), carbon monoxide (CO), hydrogen bromide (HBr), hydrogen sulfide (H_2S) and nitrogen (N_2)?

6 The red colour of blood is due to iron. Would this colour be due to iron atoms or iron ions? Explain your answer.

7 Why is it wrong to speak of *molecules* of sodium chloride? What should you say instead?

Challenge solutions

1 It would need to lose two electrons.

2

Type of bond	*Kind of atoms*	*Electrons are ...*	*Result*
covalent	non-metals	shared	molecules
ionic	metals and non-metals	lost and gained	ionic compound

3 Ionic compounds are neutral because the number of positive charges is equal to the number of negative charges. For example, sodium chloride has equal numbers of sodium ions and chloride ions.

4 Positive copper ions are never found on their own because they readily accept electrons to form neutral copper atoms or combine with negative ions to form ionic compounds.

5 You would expect molecules containing different atoms to be polar; that is, NH_3, CO, HBr and H_2S.

6 You would expect the red colour to be due to iron ions because iron (consisting of iron atoms) is grey in colour.

7 It is wrong because there are no separate molecules containing one atom of sodium and one atom of chlorine. It is more accurate to say that sodium chloride is an ionic compound consisting of many sodium and chlorine ions which form a crystal lattice.

10.2 Chemical shorthand

Chemical formulas

Covalent compounds are formed when atoms join to form molecules; and ionic compounds form when positive and negative ions join together in a crystal lattice. A shorthand way of representing a compound is to use a chemical formula made up from the symbols of the elements in it. For example, the formula for carbon dioxide is CO_2. The 2 after the oxygen shows that there are two atoms of oxygen in a molecule of carbon dioxide. (Note that the 2 is written as a subscript, a little below the line.) There is only one atom of carbon, but the 1 is never written in the formula. Similarly, the formula for water is H_2O (two atoms of hydrogen bonded to one atom of oxygen). Ammonia (NH_3) has three atoms of hydrogen bonded to one atom of nitrogen.

Sodium chloride is an ionic compound consisting of sodium and chloride ions held together by ionic bonds. Its formula is NaCl, which means that the ratio of sodium ions to chloride ions is 1:1. But how do we know that the formula is NaCl, rather than $NaCl_2$ or Na_2Cl or Na_2Cl_3? The answer is that you can tell from the combining powers of the atoms.

An atom's combining power or **valency** predicts how it will combine with other atoms through the loss, gain or sharing of electrons. For ionic compounds the valency is the same as the charge on the ion. For example, sodium atoms lose one electron to form Na^+ ions—so sodium has a valency of 1+. Oxygen atoms gain two electrons to form oxide ions O^{2-}—so oxygen has a valency of 2–.

A sodium ion Na^+ can combine with one negative ion that has a valency of 1–, eg Cl^-, to form NaCl. Sodium ions can also combine with oxide ions. However, two Na^+ ions are needed for each O^{2-} ion. Hence the formula for sodium oxide is Na_2O.

sodium Na^+ + chlorine Cl^- → sodium chloride NaCl

sodium Na^+ + oxygen O^{2-} → sodium oxide Na_2O

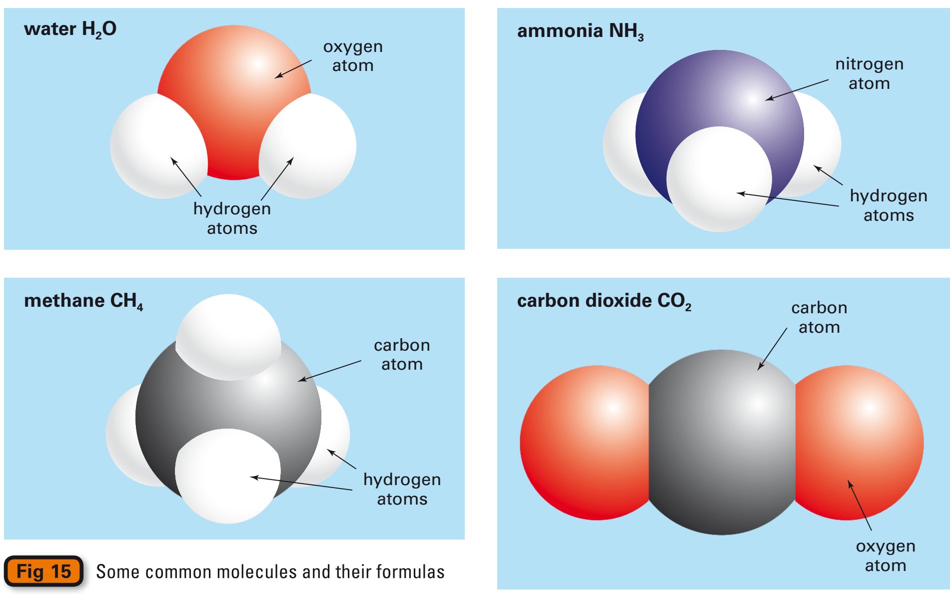


Fig 15 Some common molecules and their formulas

Hints and tips

Check that students understand the importance of the direction of the arrows in the equations. Remind them about the difference between reactants and products. It is also important students know the difference between superscript and subscript text. Make sure students write the valency showing the number first, then the positive or negative ion, eg 2+ and *not* +2.

Homework

Get students to create a 'match the word and meaning' worksheet of the chapter so far. Have them write their name on it. In the next lesson you could collect the sheets and randomly distribute them among the students. They are to complete the sheets and then have them corrected by the 'student author'. Any discrepancies should be discussed together, and then with you if necessary.

Learning experience

Ask students to skim through section 10.2 and jot down six or so questions about chemical shorthand. As they progress through this section, give students time to review and answer their questions. Any question they cannot answer can be discussed with the student next to them. If they still cannot answer the question, open it up to the whole class.

Learning experience

Does it really matter if you don't write the chemical formula properly? Demonstrate to the students why it is important to write chemical formulas correctly. Choose two substances with similar chemical formulas: water (H_2O) and hydrogen peroxide (H_2O_2) is a good example. Both chemical formulas are similar and their solutions are clear, so they look exactly the same. Does an extra oxygen atom in the hydrogen peroxide really make such a difference?

To answer this question, students may recall the hydrogen peroxide and manganese dioxide teacher demonstration from the Learning experience in *ScienceWorld 2 Teacher Edition*, page 32. To perform this demonstration:

- In a fume cabinet, measure out the same amounts of water and hydrogen peroxide solution (6%) into two large measuring cylinders—about one-fifth full.
- Add a small amount of manganese dioxide (MnO_2) to the water, and then to the hydrogen peroxide. What happens? The water does not react, but the peroxide reacts violently.
- You may like to use a glowing taper to show oxygen is present. Gloves and safety glasses are essential, and you need to take great care not to get any of the strong solution on you.

Students can write a word equation to explain what happened, then write the same equation using chemical formulas. Manganese dioxide acts as a catalyst and speeds up the reaction. H_2O and H_2O_2 are definitely *not* the same compound! What else can students remember from last year about hydrogen peroxide? (For example, it readily decomposes into water and oxygen, and in solution it is used as a bleaching agent and disinfectant—it is corrosive.)

Hints and tips

Posters are a great visual tool. Consider getting the students to make a poster displaying the information in the table on this page. Having it on permanent display will help students to familiarise themselves with chemical names, symbols and valencies.

Erratum

In earlier printings of the textbook there is an error in the second paragraph on this page, where 'two nitrate NO^{3-} compound ions' are mentioned. This should read 'two nitrate NO_3^- compound ions'.

Research

Copper sulfate and sodium chloride form crystal lattices. Do all ionic and covalent compounds form lattices when they are solids? Explain. Get students to research why some compounds form lattices, and others don't. Students could present their work as an article for a science magazine or a newspaper. They could write their information as a 'breaking news' story—as if this is the first time the discovery has been made.

Sometimes the ions consist of groups of atoms called *compound ions*. For example, copper sulfate is made up of copper ions and sulfate ions. Each copper ion (Cu^{2+}) has two positive charges. Each sulfate compound ion (SO_4^{2-}) consists of one sulfur atom covalently bonded to four oxygen atoms, as shown. The five atoms of the sulfate ion usually stay together in chemical reactions.

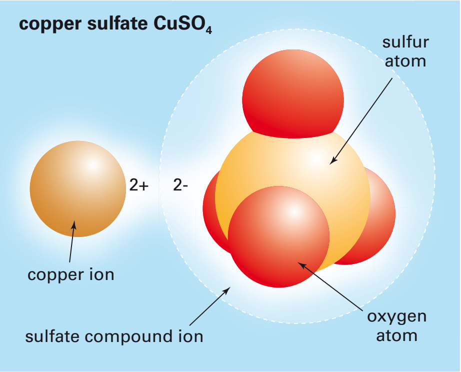


Fig 16 A model of copper sulfate $CuSO_4$. Note that these ions would be arranged in a crystal lattice similar to that for sodium chloride (Fig 6 on page 238).

The sulfate compound ion has two negative charges which balance the two positive charges on the copper ion. Similarly each unit of copper nitrate $Cu(NO_3)_2$ contains one Cu^{2+} ion and two nitrate NO_3^- compound ions; and each sodium carbonate Na_2CO_3 contains two Na^+ ions and one carbonate CO_3^{2-} compound ion.

The valencies of various ions are listed in the table on the right. There is a pattern in the table which will make sense when you study the next chapter. Note that some elements have two different valencies. For example, copper loses two electrons to form Cu^{2+} ions; but sometimes it loses only one electron to form Cu^+ ions. Because of this it forms two different compounds with oxygen. To distinguish between these compounds you include the valency of the copper ion in the formula, using Roman numerals.

	Ion	Symbol	Valency
	ammonium	NH_4^+	1+
	hydrogen	H^+	1+
	potassium	K^+	1+
	silver	Ag^+	1+
	sodium	Na^+	1+
Metals	calcium	Ca^{2+}	2+
	copper	Cu^{2+}	2+ (or 1+)
	lead	Pb^{2+}	2+ (or 1+)
	magnesium	Mg^{2+}	2+
	zinc	Zn^{2+}	2+
	aluminium	Al^{3+}	3+
	iron	Fe^{3+}	3+ (or 2+)
Non-metals	bromide	Br^-	1–
	chloride	Cl^-	1–
	hydrogen carbonate	HCO_3^-	1–
	hydroxide	OH^-	1–
	iodide	I^-	1–
	nitrate	NO_3^-	1–
	carbonate	CO_3^{2-}	2–
	oxide	O^{2-}	2–
	sulfate	SO_4^{2-}	2–
	sulfide	S^{2-}	2–
	sulfite	SO_3^{2-}	2–
	phosphate	PO_4^{3-}	3–

This means the copper has a valency of 1.

copper(I) oxide Cu_2O
copper(II) oxide CuO

This means the copper has a valency of 2.

Learning experience

Students could make their own plasticine models of ions from the list of 24 ions in the table. Decide on a colour code and approximate atom size for each element. Each student could be assigned an ion. They should label it as a positive or negative ion, and a metal or non-metal. They should also indicate, using the table, which types of ions it is likely to bond with. Toothpicks or skewers could be used to put signs on the ions showing the valency. If there are more than 24 students in the class, the different valencies for copper, lead and iron could be shown. Students could then write their own fact sheet about their ion. Models and fact sheets could be displayed in a cabinet in the science complex or classroom.

Writing ionic formulas

To write the formula for an ionic compound, follow these rules.

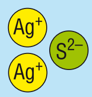
Example 1

1 Write down the symbols of the ions. Note that the positive ion (usually a metal) goes first. Write the valencies above the symbols.

silver   sulfide
1+   2–
Ag   S

2 Crisscross the valencies to get correct subscripts. Leave out the + and – signs.

1+   2–
Ag 2   S 1

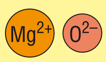


Ag_2S

3 Write the correct formula with subscripts, leaving out the 1. Note that the charges are balanced. You need two Ag^+ ions to balance one S^{2-} ion.

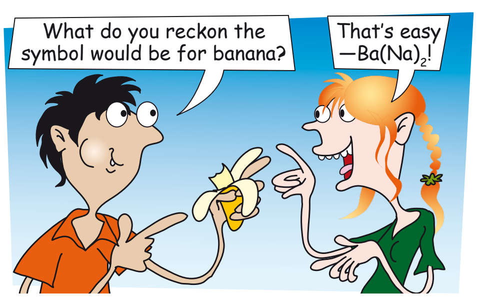
Example 2

magnesium   oxide
2+   2–
Mg 2   O 2



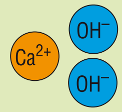
MgO

In this example you simplify Mg_2O_2 to MgO since the magnesium and oxygen are in the ratio 1:1.



Example 3

calcium   hydroxide
2+   1–
Ca 1   (OH) 2

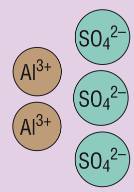


$Ca(OH)_2$

In this example you put the compound ion in brackets to indicate that it acts as a single unit. The subscript 2 refers to everything inside the brackets. In other words, there are two oxygen atoms and two hydrogen atoms. Where there is only one unit of the compound ion you can drop the brackets, eg NaOH rather than Na(OH).

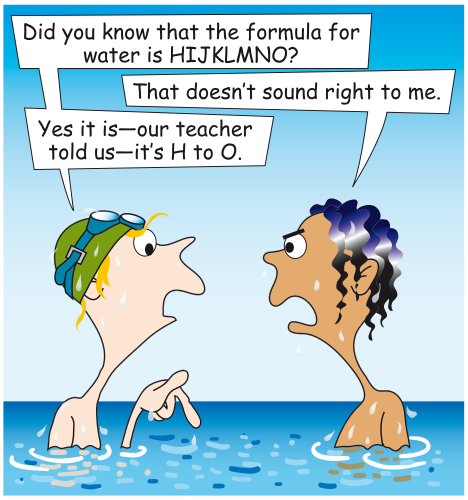
Example 4

aluminium   sulfate
+3   –2
Al 2 (SO_4) 3



$Al_2(SO_4)_3$

Two 3+ charges balance three 2– charges.



Hints and tips

- Revise the chapter by asking the students to write a paragraph about what they have learned so far and share it with the person next to them. Then choose a few students to read their paragraphs aloud to the class. It is always good for students to realise how much they have learned.
- Make sure that you monitor students who experience difficulties with maths as they might find the concept of writing formulas challenging. For this section, consider pairing students according to their maths and science abilities, so that students who experience difficulties can be assisted by those who are more able.

Learning experience

Ionic bonding game

Assign each student an ion from the table on page 244. On A5 pieces of card they are to write out the ion's name and its symbol. They will need to make three identical cards (apart from trivalent ions, where only two cards are needed). Have the students string the cards around their necks so that they can easily be read. Now get students to pair up by making ionic bonds. Remind them that positives and negatives attract. Each pair needs to work out how many ions are needed so that the compound is balanced. For example, calcium hydroxide needs one Ca^{2+} ion and two OH^- ions to make $Ca(OH)_2$, while aluminium sulfate needs two Al^{3+} ions and three SO_4^{2-} ions. Any extra ions need to be removed.

When students have paired up, have them write down the balanced compounds in their notebooks, then pair up with another 'student ion'. See how many compounds students can balance in a set time. This activity is a great revision tool and can be used at the beginning or end of chemistry lessons. It can be played more than once, and will be useful in the next chapter. Make sure you collect the cards at the end of the game.

Activity notes

- This is an excellent activity for use in class or at home.
- The simplest way is to make six copies of page 246, cut out the templates and paste them onto another sheet of paper. Then make one copy for each student.
- Another way is to make copies of each ion on different-coloured cards, cut these out and keep them in plastic bags so that they can be reused by other classes. Students then trace around the shapes rather than paste them into their books. It is a good idea to laminate the cards.

Erratum

In earlier printings of the textbook, there is an error in the labelling of the ions for the cards. The second, fifth and sixth cards should read Mg^{2+}, CO_3^{2-} and PO_4^{3-}.

Hints and tips

- Revise the difference between ionic and covalent bonding. Remind students that in covalent bonding, the atoms share electrons.
- It might be helpful to show students the following for the example given showing carbon and oxide forming carbon dioxide:

carbon oxide
4 2

C_2 O_4

C_2O_4 The 2:4 ratio is simplified to 1:2 and written as CO_2.
↓
CO_2

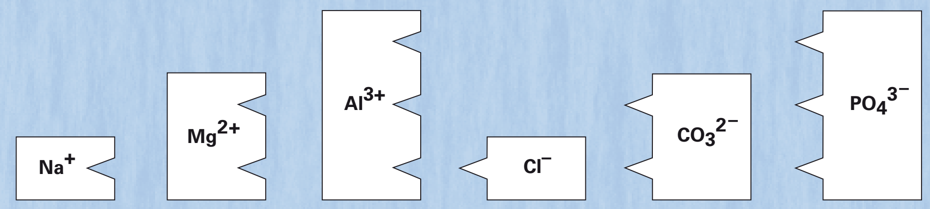
- Show students another example of using valencies to write the formula of a covalent compound.

sulfur oxide
6 2

S_2 O_6

S_2O_6 The 2:6 ratio is simplified to 1:3 and written as SO_3.
↓
SO_3

Activity



You can use paper or cardboard cut-outs to make models of various compounds. Look at the models above for various ions. Notches represent electrons lost to form positive ions. One notch indicates one positive charge or a valency of 1, two notches indicate a valency of 2, and three notches a valency of 3. Similarly, spikes represent electrons gained to form negative ions. Work in a group so you can share the work.

1. Your teacher will give you copies of the models—six of each. Colour each type of ion a different colour, eg sodium ions yellow, chloride ions green etc.
2. Use scissors to cut out the models.
3. Use your cut-outs to make a model of sodium chloride. Glue the model compound into your notebook and write its formula under it.
4. Make models of the following compounds:
   sodium carbonate
   sodium phosphate
   magnesium chloride
   magnesium carbonate
   magnesium phosphate
   aluminium chloride
   aluminium carbonate
   aluminium phosphate
5. For each compound you make, count the number of positive charges and the number of negative charges to make sure they are equal. Then write the formula under the model.

Covalent compounds

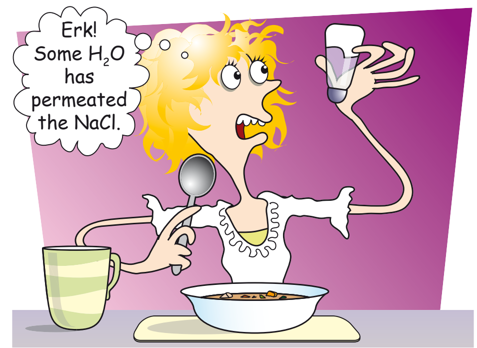
You can also use valencies to write the formulas for covalent compounds. For example:

carbon oxide
4 2
C 2 O 4

CO_2

Instead of calling this compound carbon oxide, it is usually called carbon dioxide. The *di* comes from the Greek word for 'two' since two oxygen atoms bond to one carbon atom. Similarly, CO is carbon monoxide, and SO_3 is sulfur trioxide.

Some elements form molecules containing a pair of atoms. These are called *diatomic* molecules, eg hydrogen H_2, oxygen O_2, nitrogen N_2, chlorine Cl_2 and iodine I_2. It is best to remember these formulas.



Learning experience

Ask students to construct a table listing the prefixes mono–, di– and tri–, defining the meaning of each one, and giving an example in the context of chemistry. Then ask students to try to write down what they think the formulas of some compounds are using their tabulated information. For example, what do they think dihydrogen monoxide or dicarbon dihydride is?

Chemical equations

A chemical equation is a shorthand way of writing down what happens in a chemical reaction. To write an equation you must know the names of the reactants and the products. If you don't know these you cannot write the equation. Sometimes, however, you can predict what the products might be, if you know the reactants.

Work through the following three examples step by step.

Equation 1

Zinc metal reacts with hydrochloric acid to produce hydrogen gas and a solution of an ionic compound called zinc chloride.

Step 1: Writing the word equation

zinc + hydrochloric acid → zinc chloride + hydrogen

Step 2: Writing down the formulas

You next write down the formulas of the reactants and products. For most elements the atoms exist on their own, so you simply write the symbol for the element, in this case zinc Zn. It is best to remember that the formula for hydrochloric acid is HCl. For zinc chloride you can work it out using valencies from the table on page 244. In hydrogen gas, however, the molecules are diatomic (H_2).

$$Zn + HCl \rightarrow ZnCl_2 + H_2$$

Step 3: Balancing the equation

In a chemical reaction the atoms are rearranged, but you end up with the same number of atoms as you started with. (This is the law of conservation of mass.) So the final step in writing an equation is to make sure that the numbers of atoms of each element are the same on both sides of the equation. This is called *balancing the equation.*

In this case, there is one zinc atom on each side of the equation. So the zinc atoms are balanced. There are two hydrogen atoms on the right-hand side, but only one on the left. So you can balance the hydrogen atoms by putting a 2 in front of the HCl on the left-hand side. This means 2 molecules of HCl.

$$Zn + 2HCl \rightarrow ZnCl_2 + H_2$$

The 2 refers to every atom in the formula. So 2HCl means that there are two atoms of H and two atoms of Cl. So the chlorine atoms are balanced. *Never change the number in a formula to balance an equation.* The balancing numbers always go in front of the formulas.

Symbols are usually added to tell you whether the reactants are solids (s), liquids (l), gases (g) or dissolved in water (aq—short for aqueous). For example:

$$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

$Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

Hints and tips

- Remind students that the conservation of mass applies to chemical reactions. There is the same number of atoms before the reaction as there is after the reaction. That is, the number of atoms of the reactant(s) equals the number of atoms of the product(s).
- Explain to the students that (s), (l), (aq) and (g) represent the states of matter—solid, liquid, aqueous solution, and gas.

Learning experience

Students could make their own balancing equations worksheet. Ask them to divide an A4 page down the middle, and write unbalanced equations on one side and balanced equations on the other. The page is to be folded in half so that only the unbalanced equations are displayed. In pairs, students can swap sheets and balance each other's equations. Alternatively, the sheets could be collected and photocopied, then given as revision questions for the next few lessons. You might like to keep some sheets for Chapter 11.

Learning experience

Devise a game in which students have to see how many questions they can answer correctly in a given time frame. Prepare a series of cards with questions written on one side and a number written on the other. About 50 cards should be sufficient. Position the cards around the room—you could hide some to make the game more interesting. Give each student a worksheet with a list of question numbers and adequate space to fill in their answer. When they think they have correctly answered a question, they present it to you for checking. Students can only move on to a new card if they correctly answer the question, and must return each card to its original position in the room before attempting a new question. It is best to sit at the front of the room so you can survey the class and students can easily find you to check their answers. Have a master copy of the solutions already prepared for easy marking.

Include a variety of questions such as balancing equations, stating if a compound is ionic or covalent, calculating the number of atoms on the reactant or product side of an equation, writing word equations, and so on. This task works well individually or in pairs. Consider awarding those students with the greatest number of correct answers some small prizes, such as fancy pens, stickers or food treats.

Hints and tips

- For more able students, prepare some worksheets on balancing tricky equations. Combine both the chemical symbols and chemical names. The science department might already have some blackline masters or other resources available for photocopying, which may save time. Consider having a selection of worksheets graded according to their level of difficulty. Gifted and talented students should be encouraged to attempt the most difficult level.
- Remember to monitor ESL students and check they understand words and their meanings. Suggest they translate the names of the given chemical compounds in the examples into their native language.

Equation 2

Ammonia is a very important gas used in industry to make nitric acid, fertilisers, drugs, dyes and plastics. It is made by reacting the gases nitrogen and hydrogen at a high temperature in the presence of a catalyst.

nitrogen + hydrogen → ammonia

You need to know the formulas for nitrogen, hydrogen and ammonia.

$$N_2 + H_2 \rightarrow NH_3$$

To balance the nitrogen atoms you have to add a 2 in front of NH_3.

$$N_2 + H_2 \rightarrow 2NH_3$$

This makes 6 atoms of hydrogen on the right-hand side. So to balance the hydrogens you have to add a 3 in front of H_2 on the left-hand side.

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

As a final check on the balancing you can write down the numbers of atoms on each side of the equation, as shown.

LEFT		RIGHT	
N	2	N	2
H	3 × 2 = 6	H	2 × 3 = 6

Equation 3

Iron rusts when it reacts with oxygen in the air to produce iron(III) oxide.

iron + oxygen → iron oxide

$$Fe + O_2 \rightarrow Fe_2O_3$$

To balance the iron atoms, put a 2 in front of Fe.

$$2Fe + O_2 \rightarrow Fe_2O_3$$

To balance the oxygen atoms you need a 3 in front of O_2 and a 2 in front of Fe_2O_3. You may need to work this out by trial and error.

$$2Fe + 3O_2 \rightarrow 2Fe_2O_3$$

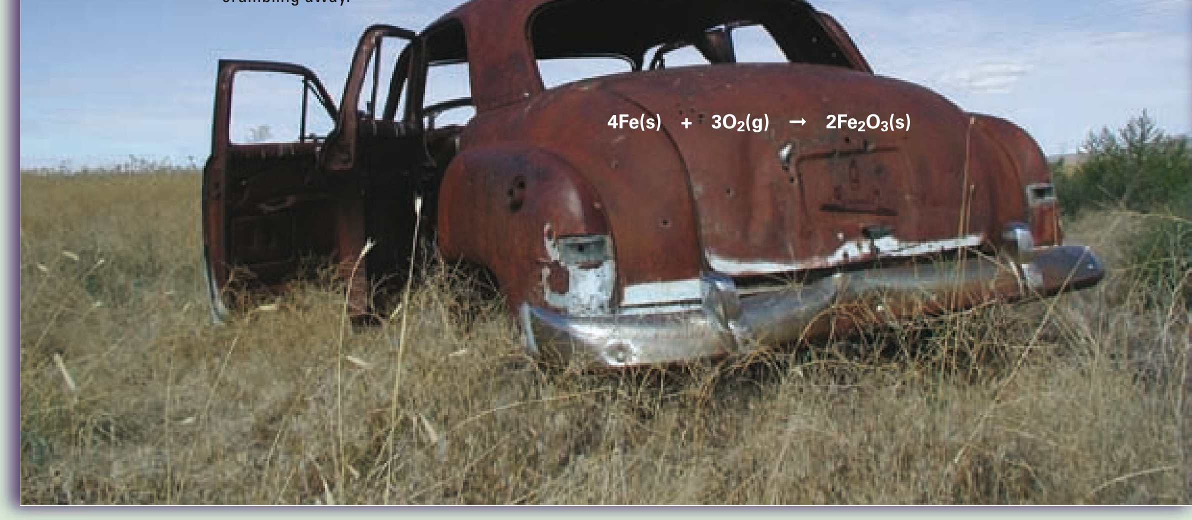
Finally, you need to balance the iron atoms again.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

LEFT		RIGHT	
Fe	4	Fe	2 × 2 = 4
O	3 × 2 = 6	H	2 × 3 = 6

Note that the balancing numbers you use should be the smallest possible. For instance, the equation $8Fe + 6O_2 \rightarrow 4Fe_2O_3$ is balanced, but can be simplified to $4Fe + 3O_2 \rightarrow 2Fe_2O_3$.

Fig 22 The iron in this old car is slowly turning to rust (iron oxide) and crumbling away.



Learning experience

A good revision activity is for the students to come up with their own song or rap about balancing chemical equations. This activity can be tackled individually or in small groups, and it is best to set a time limit. You might like to suggest they write their song to a familiar tune such as 'Mary Had a Little Lamb', 'Money, Money, Money' or 'Advance Australia Fair'. You could start them off with something like:

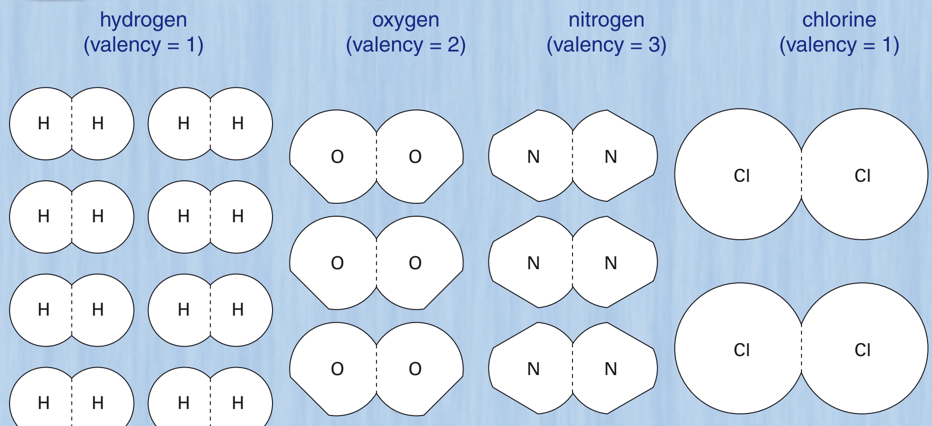
- 'Chemistry, Chemistry, Chemistry in a formula world …' to the tune of 'Money, Money, Money'
- 'Metal ions all let us rejoice, as valence electrons free …' to the tune of 'Advance Australia Fair'.

Students might like to pre-record their song as a video instead of performing it in front of the class.

Learning experience

Ask students to prepare a multimedia presentation (eg PowerPoint) showing balancing equations. The presentations need to show the steps in how an equation is balanced. Give each student a different unbalanced equation to use for their presentation. They need to identify the reactant(s) and the product(s), and write down the word equation and the steps involved to make it balanced. Encourage students to use the animation features of PowerPoint and coloured fonts to enhance their work. Pages 250–251 have examples of equations that could be given to students for this task.

Activity



You can use paper or cardboard cut-outs to make models of chemical reactions. The cutouts above represent some diatomic molecules.

1. Photocopy or trace the molecular models above and label the atoms. Colour the atoms different colours, eg white for hydrogen, red for oxygen, blue for nitrogen and green for chlorine.
2. The example below shows how the models can be used to represent the reaction between hydrogen and oxygen to form water.
3. Use the models to represent the following reactions:

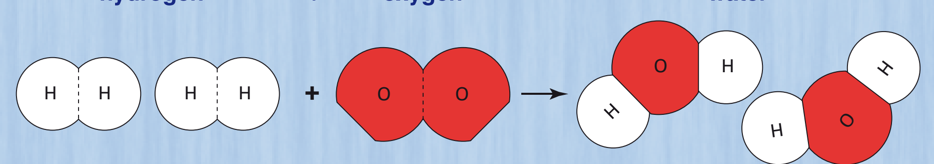
   hydrogen + chlorine → hydrogen chloride

   nitrogen + hydrogen → ammonia

   nitrogen + oxygen → nitrogen dioxide

   You will need to break the molecules into their atoms by cutting along the dotted lines. Unless you do this the 'reaction' cannot occur.
4. Glue the models into your notebook and write the balanced equations under them.
5. If you have a molecular models kit you can make three-dimensional models.

hydrogen + oxygen → water

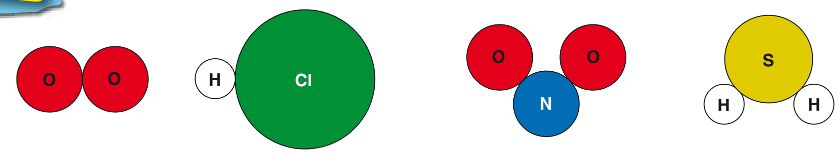


$$2H_2 + O_2 \rightarrow 2H_2O$$

Activity notes

- As for the previous activity (page 246), photocopy these onto sheets to reduce paper wastage. Alternatively, students could trace the cut-outs onto scrap paper from the recycled paper box.
- Small plasticine models could be made, or wooden beads glued together.

Check!



1 Look at the models of the four molecules above, and write the chemical formula for each of them.

2 Sulfuric acid contains hydrogen, sulfur and oxygen in the ratio 2:1:4. What is its formula?

3 One of the substances in superphosphate fertiliser has the formula $Ca(H_2PO_4)_2$. How many atoms of calcium, hydrogen, phosphorus and oxygen are represented by this formula?

4 Use the table on page 244 to write the correct formula (with its electric charge) for each of the following compound ions:
- **a** ammonium
- **b** carbonate
- **c** hydrogen carbonate
- **d** hydroxide
- **e** nitrate
- **f** phosphate
- **g** sulfate
- **h** sulfite

5 Name the following compounds:
- **a** KCl
- **b** $NaNO_3$
- **c** MgO
- **d** $FeCl_3$
- **e** $Al(OH)_3$
- **f** $(NH_4)_3PO_4$
- **g** HCl
- **h** $NaHCO_3$

6 Copy this table and complete the formulas.

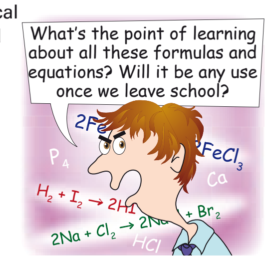
	chloride	sulfate	phosphate
calcium			
iron(III)			
sodium			

7 Write down the formulas of the following compounds:
- **a** sodium hydroxide
- **b** ammonium sulfate
- **c** hydrogen sulfide
- **d** sodium sulfite
- **e** calcium hydrogen carbonate
- **f** magnesium sulfate
- **g** calcium phosphate
- **h** copper(II) hydroxide
- **i** iron(II) oxide

8 Carlos heated some blue copper(II) nitrate in a test tube. A brown gas (nitrogen dioxide) and a colourless gas (oxygen) were produced and black copper(II) oxide was left in the test tube.
- **a** List the reactants and products.
- **b** Write a word equation for the reaction.

9 Copy these equations into your notebook and balance them where necessary. The formulas are all correct and must not be changed.
- **a** $NaBr + Cl_2 \rightarrow NaCl + Br_2$
- **b** $Fe + Cl_2 \rightarrow FeCl_3$
- **c** $Ba(NO_3)_2 + H_2SO_4 \rightarrow BaSO_4 + HNO_3$
- **d** $CaCO_3 \rightarrow CaO + CO_2$
- **e** $H_2 + I_2 \rightarrow HI$
- **f** $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
- **g** $NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$
- **h** $P_4 + O_2 \rightarrow P_2O_5$
- **i** $CH_4 + Cl_2 \rightarrow CCl_4 + HCl$
- **j** $Pb(NO_3)_2 \rightarrow PbO + NO_2 + O_2$

10 Michael can't see the point of learning about chemical formulas and equations. In a group discuss how you could answer him.



Check! solutions

1 The formulas (from left to right) are: O_2, HCl, NO_2, H_2S.

2 The formula is H_2SO_4.

3 In each molecule there is one calcium atom, four hydrogen atoms, two phosphorus atoms and eight oxygen atoms.

4

	Formula	*Charge*
a	NH_4	1+
b	CO_3	2–
c	HCO_3	1–
d	OH	1–
e	O	2–
f	PO_4	3–
g	SO_4	2–
h	SO_3	1–

5 The names of these compounds are:
- **a** potassium chloride
- **b** sodium nitrate
- **c** magnesium oxide
- **d** iron(III) chloride
- **e** aluminium hydroxide
- **f** ammonium phosphate
- **g** hydrogen chloride
- **h** sodium hydrogen carbonate.

6

	chloride	*sulfate*	*phosphate*
calcium	$CaCl_2$	$CaSO_4$	$Ca_3(PO_4)_2$
iron(III)	$FeCl_3$	$Fe_2(SO_4)_3$	$FePO_4$
sodium	NaCl	Na_2SO_4	Na_3PO_4

7
- **a** NaOH
- **b** $(NH_4)_2SO_4$
- **c** H_2S
- **d** Na_2SO_3
- **e** $Ca(HCO_3)_2$
- **f** $MgSO_4$
- **g** $Ca_3(PO_4)_2$
- **h** $Cu(OH)_2$
- **i** FeO

8 In this reaction:
- **a** The reactant is copper(II) nitrate. The products are nitrogen dioxide, oxygen and copper(II) oxide.
- **b** copper nitrate → nitrogen dioxide + oxygen + copper oxide

9 The balanced equations are:
- **a** $2NaBr + Cl_2 \rightarrow 2NaCl + Br_2$
- **b** $2Fe + 3Cl_2 \rightarrow 2FeCl_3$
- **c** $Ba(NO_3)_2 + H_2SO_4 \rightarrow BaSO_4 + 2HNO_3$
- **d** $CaCO_3 \rightarrow CaO + CO_2$
- **e** $H_2 + I_2 \rightarrow 2HI$
- **f** $Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$
- **g** $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$
- **h** $P_4 + 5O_2 \rightarrow 2P_2O_5$
- **i** $CH_4 + 4Cl_2 \rightarrow CCl_4 + 4HCl$
- **j** $2Pb(NO_3)_2 \rightarrow 2PbO + 4NO_2 + O_2$

10 Michael should be helped to realise that we are surrounded by chemicals. Chemical names and formulas are often seen on food packaging, medicines, laundry products, and in the shed on fertilisers and garden sprays. It is important for him to know how to handle these chemicals safely and how to recognise dangerous chemicals. Michael is only in Year 10 and probably has no clear idea of his career. It is possible that he might one day have a job that will require a detailed knowledge of chemicals.

challenge

1 Look at this balanced equation:
$Mg + O_2 \rightarrow 2MgO$
   a How many atoms of magnesium react with one molecule of oxygen?
   b How many atoms of magnesium react with 5 million molecules of oxygen?

2 Each of the following equations is incorrect. Rewrite them correctly and balance them.
   a $H + Cl \rightarrow HCl$
   b $H_2O_2 \rightarrow H_2O + O$
   c $Cu + O_2 \rightarrow CuO_2$
   d $Pb(NO_3)_2 + KI \rightarrow PbI + K(NO_3)_2$
   e $Na + H_2O \rightarrow NaOH + H$

3 Write balanced equations for the following reactions.
   a Zinc reacts with hydrochloric acid to produce zinc chloride and hydrogen.
   b Sulfur dioxide (SO_2) burns in oxygen to produce sulfur trioxide (SO_3).
   c Copper(II) carbonate reacts with hydrochloric acid to produce copper(II) chloride, water and carbon dioxide.
   d During cooking, sodium hydrogen carbonate (baking soda) decomposes to sodium carbonate, carbon dioxide and oxygen.
   e When heated, cane sugar ($C_{12}H_{22}O_{11}$) decomposes to give carbon and water.
   f During respiration, glucose ($C_6H_{12}O_6$) reacts with oxygen to give carbon dioxide and water.

4 Elements X and Y form compounds with carbon with the formulas CX_4 and CY_2. Predict the formula of a compound of X and Y. Explain how you worked it out.

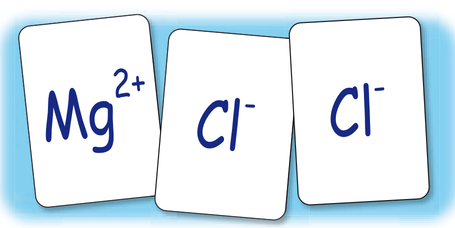
5 A metal M forms a sulfate with the formula $M_2(SO_4)_3$. Given this information, which of the following formulas are correct?
   a M_2O
   b $M_2(CO_3)_3$
   c $M(OH)_3$
   d M_2Cl_3
   e MPO_4
   f M_3S_2

6 Challenge yourself by trying to balance these three equations:
   a $Al(NO_3)_3 + K_2Cr_2O_7 \rightarrow Al_2(Cr_2O_7)_3 + KNO_3$
   b $FeCl_2 + HNO_3 + HCl \rightarrow FeCl_3 + NO + H_2O$
   c $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + H_2O + NO$

try this

For this activity you will need playing cards with the names and valencies of positive and negative ions. Use all the ions listed on page 244 and make four identical cards for each one. You will need a total of 96 cards. (Use the 2+ valency for Cu, Fe and Pb.)

Play the game in a group of four or five. The aim of the game is to combine cards to make compounds with correct formulas; for example, one magnesium Mg^{2+} card goes with two chloride Cl^- cards to give $MgCl_2$.



The dealer shuffles the cards and gives each player seven cards. When it is your turn you try to make a compound using the cards in your hand. If you use two cards to make the compound then you pick up two cards to replace them, and so on. If you can't make a compound you pick up another card. You can also choose to pick up another card instead of making a compound.

The game continues until all the cards are used up. The winner is the player who has made the most compounds with the correct formula.

Challenge solutions

1 a Two atoms of magnesium react with each molecule of oxygen.
   b 10 million atoms of magnesium react with 5 million molecules of oxygen (same ratio).

2 These are the correct equations:
   a $H_2 + Cl_2 \rightarrow 2HCl$
   b $2H_2O_2 \rightarrow 2H_2O + O_2$
   c $2Cu(II) + O_2 \rightarrow 2CuO$
   or
   $4Cu(I) + O_2 \rightarrow 2Cu_2O$
   d $Pb(NO_3)_2 + 2KI \rightarrow PbI_2 + 2KNO_3$
   e $2Na + 2H_2O \rightarrow 2NaOH + H_2$

3 The balanced equations are:
   a $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
   b $2SO_2 + O_2 \rightarrow 2SO_3$
   c $CuCO_3 + 2HCl \rightarrow CuCl_2 + H_2O + CO_2$
   d $2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$
   e $C_{12}H_{22}O_{11} \rightarrow 12C + 11H_2O$
   f $C_6H_{12}O_6 + 6O_2 \rightarrow 6H_2O + 6CO_2$

4 The valency of carbon is 4; hence, the valency of X must be 1 and the valency of Y must be 2. So you would expect the formula to be X_2Y.

5 The valency of metal M must be 3+
The correct formulas are **b**, **c** and **e**.

6 The correct equations are:
   a $2Al(NO_3)_3 + 3K_2Cr_2O_7 \rightarrow Al_2(Cr_2O_7)_3 + 6KNO_3$
   b $3FeCl_2 + HNO_3 + 3HCl \rightarrow 3FeCl_3 + NO + 2H_2O$
   c $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 4H_2O + 2NO$

Hints and tips

For the Try this activity, cards may be available from science suppliers or office supply stores (blank business cards). Otherwise they can be made from card using a guillotine.

Research

Students could research types of chemical reactions and make a booklet explaining each. The four basic types of chemical reactions are combination, precipitation, decomposition (breaking-down) and combustion reactions. Some students may choose to present their booklet electronically. Encourage accuracy, simple language and colourful, well-labelled diagrams. Gifted and talented students could make their presentation interactive.

10.3 Predicting a reaction

In a chemical reaction substances react to form new substances. But what happens to the atoms, molecules and ions in the reactants? To answer this, consider the reaction between potassium iodide and lead nitrate solutions.

Potassium iodide is an ionic compound with the formula KI. It consists of K^+ ions and I^- ions. In solid potassium iodide, the ions are packed closely in a crystal lattice, like the sodium chloride lattice (Fig 6 on page 238). When the solid dissolves in water, the positive and negative ions break apart and spread throughout the solution.

$$KI \rightarrow K^+ + I^-$$

Similarly, lead nitrate $Pb(NO_3)_2$ breaks up to form lead ions Pb^{2+} and nitrate ions NO_3^- in solution.

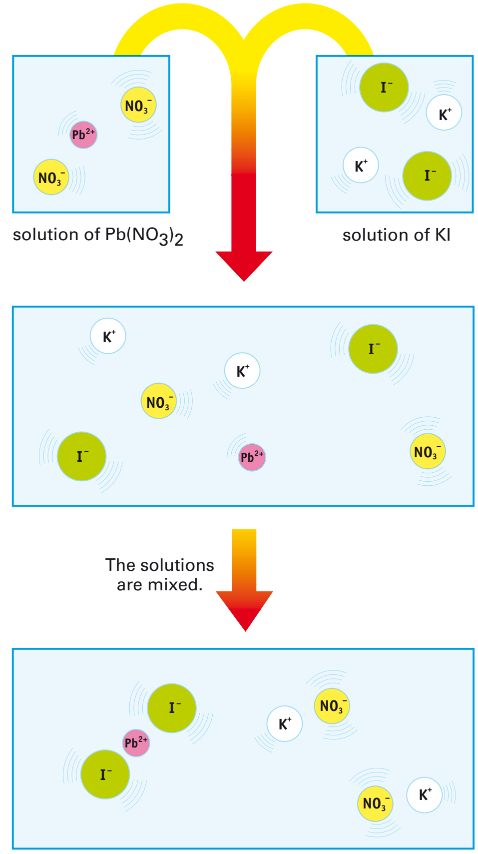
$$Pb(NO_3)_2 \rightarrow Pb^{2+} + 2NO_3^-$$

You can now try to predict what might happen when potassium iodide and lead nitrate solutions react. You can use spheres of different sizes and colours, as shown on the right, to represent the various ions.

When the two solutions are mixed, the four ions also mix. (This is like putting all the spheres in a box and shaking it.) The ions are constantly moving, and there is a good chance that they will bump into each other and possibly combine.

Lead ions can bump into nitrate ions or iodide ions, but are not likely to bump into potassium ions. This is because lead ions and potassium ions both have the same charge, and like charges repel. Similarly, potassium ions can bump into nitrate ions and iodide ions.

So, there is a possibility that the ions could change partners. Lead ions could combine with iodide ions to form lead iodide. Similarly, potassium ions could combine with nitrate ions to form potassium nitrate.



Prediction: the ions can change partners and form new substances.

So, you can predict that when lead nitrate and potassium iodide are mixed, two new substances (lead iodide and potassium nitrate) could be formed. The word equation for the reaction is shown below. You can test this prediction in Investigate 20.

lead nitrate + potassium iodide → lead iodide + potassium nitrate

Investigate

20 WHAT IS THE PRECIPITATE?

Aim

To investigate the precipitate that forms when lead nitrate and potassium iodide solutions are mixed.

PART A
Forming the precipitate

Materials

- 2 test tubes
- test tube rack
- **lead nitrate** solution (0.1M)
- **potassium iodide** solution (0.1M)
- piece of filter paper
- filter funnel
- stand and ring clamp
- small beaker
- wash bottle containing water
- watch glass
- disposable gloves (optional)

Toxic

Wear safety glasses.

Planning and Safety Check

Read both parts of the investigation carefully.

- What is the aim of Part B?

Lead and all lead compounds are toxic. So be very careful not to get them on your hands or clothing. **Wash your hands well after using these chemicals.** You must follow the method and safety precautions closely.

- How will you dispose of the lead and lead compounds?

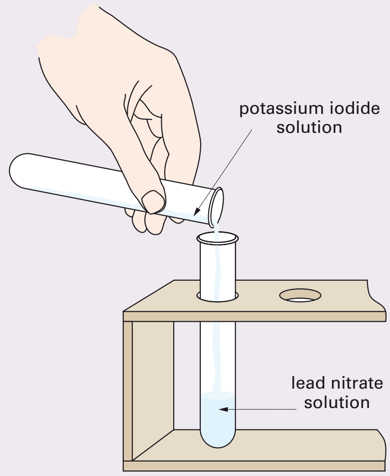
Since iodine fumes are produced in Part B, Step 2 should be done in a fume cupboard.

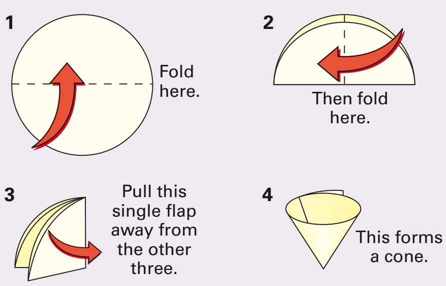
Method

1 Quarter fill one test tube with lead nitrate solution, and the other with potassium iodide solution.

2 Pour the potassium iodide solution into the lead nitrate solution.

Describe in your own words what happens.





3 Fold a filter paper and open it out into a cone, as shown. You can also use fluted filter paper.

4 Place the cone into the filter funnel and use the wash bottle to wet the paper to hold it in place. Carefully pour the contents of the test tube into the funnel and collect the filtrate in the beaker. Pour this into a special waste container.

Lab notes

- This investigation is best completed over two consecutive days.
- Make sure appropriate safety measures are taken when using the lead nitrate and potassium iodide.
- To be environmentally responsible, it is important to minimise the amounts of chemicals used in this experiment.
- Disposal of the products should be done according to local requirements and procedures.

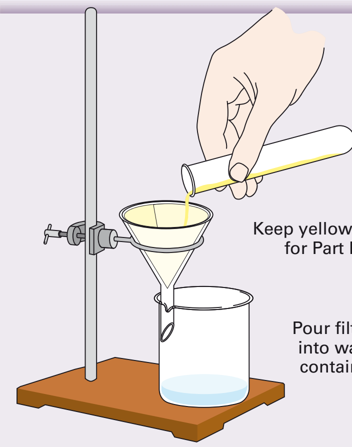
Part A

- Use bench mats throughout to reduce the risk of staining benches and books.
- Students should write their initials on the filter paper during drying to avoid confusion. It is best to use pencils.

Lab notes

Part B

- You may choose to do this part as a teacher demonstration, otherwise be vigilant in checking that all students behave responsibly.
- Ensure good ventilation when heating the solid in the flame.
- Students should wear disposable gloves for this part.
- Barbecue matches or narrow wooden splints could be used instead of the normal length matches to reduce the risk of students burning their fingers.
- A hand lens is useful to see the small balls of lead on the match.



5 Lift out the paper cone, spread it out on a watch glass and leave the yellow solid to dry, overnight if possible. You may be able to use a drying oven.
Keep the yellow solid for Part B where you will test to see which ions it contains.

Discussion

1 The yellow solid is called a precipitate. Could it be lead nitrate or potassium iodide? Explain.

2 According to the equation at the bottom of page 252, which two substances could the precipitate be?

3 How could you test that the filtrate contains something besides water?

PART B
Testing the precipitate

Materials

Toxic

- **yellow solid** from Part A
- 2 small test tubes
- test tube holder
- spatula (narrow type)
- burner and heatproof mat
- matches
- cotton wool
- white tile

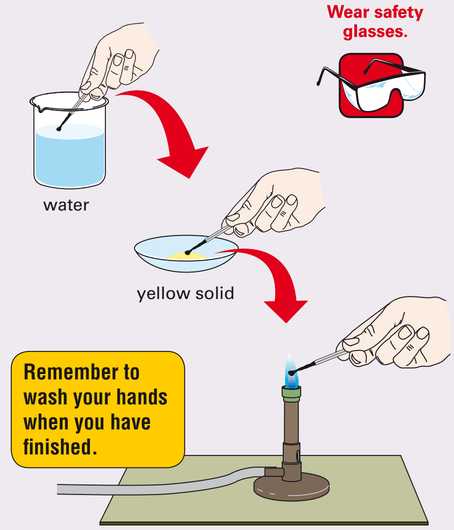
Method

1 Light a burner in the fume cupboard and adjust it so that you have a very small blue flame.

2 Use the spatula to put about half of the yellow solid into a test tube. Place a loose plug of cotton wool in the mouth of the tube (to prevent iodine fumes from escaping). Heat the test tube over the burner flame. If the solid produces purple iodine gas it contains iodide ions.

3 Light a match and let it burn briefly. The black, charred end is mainly carbon.

4 Dip the charred end into water, then into some of the yellow solid. (The water is to help the yellow solid stick to the match.)



5 Hold the tip of the match just in the burner flame for a few seconds. Watch carefully what happens. If tiny drops of silvery lead form, then the yellow solid contains lead ions. Put the match on a white tile and look for any signs of lead. A simple test is to rub it on paper. Lead leaves a black mark.

6 Place any remaining yellow solid (and the test tube from Step 2) in a special waste container.

Discussion

1 Which ions does the yellow solid contain?
2 Suggest a name for the yellow solid. Write down its chemical formula.
3 Write a balanced equation for the reaction that occurred in Part A.
4 The yellow solid was once used as a paint pigment, but it is not used any more. Suggest a reason for this.
5 When you heated lead iodide in Part B, the products were iodine gas (I_2) and lead. Write a balanced equation for this reaction.
6 When you heated lead iodide on a charred match, it reacted with oxygen in the air to form lead(II) oxide and iodine gas. The lead oxide then reacted with the carbon in the match to form lead and carbon dioxide. Write balanced equations for these two reactions.

Explaining the reaction

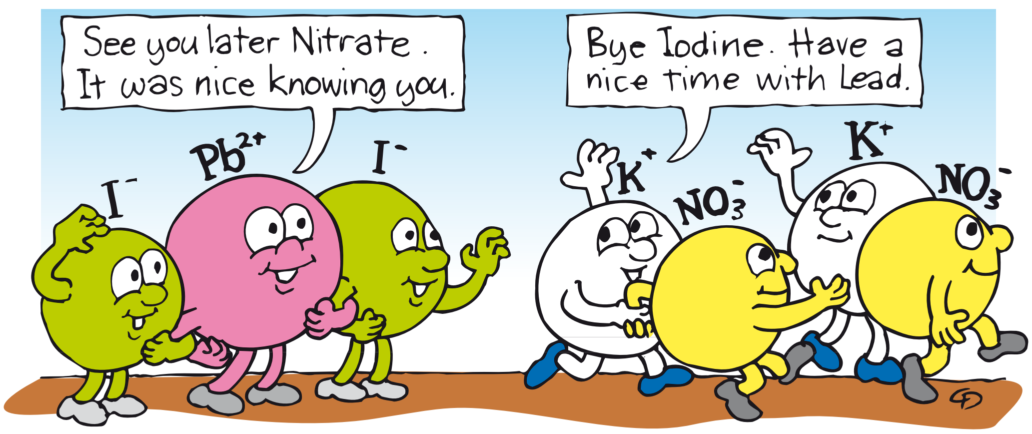
When lead nitrate and potassium iodide are mixed there are four different ions in the solution. See the diagram on page 252. The reason the lead ions combine with the iodide ions is that lead iodide is very insoluble in water. Each lead ion Pb^{2+} combines with two iodide ions I^- to form a precipitate of lead iodide PbI_2. Hence the equation for the reaction is as follows. (Check the equation you wrote in Discussion question 3 above.)

$$Pb(NO_3)_2\,(aq) + 2KI\,(aq) \rightarrow PbI_2\,(s) + 2KNO_3\,(aq)$$

This equation can be written more simply by showing only the ions which form the precipitate. This is called an *ionic equation*. Because the K^+ and NO_3^- ions do not take part in the reaction they are called *spectator ions*.

$$Pb^{2+} + 2I^- \rightarrow PbI_2$$

Fig 33 When lead nitrate and potassium iodide are mixed, the lead ions are more strongly attracted to iodide ions than they are to nitrate ions. So the ions change partners.



To see an animation of the formation of lead iodide, open **Changing partners** on the CD. Working with technology

Hints and tips

- A solid that forms when two solutions are mixed is called a precipitate. Precipitates are insoluble, that is, they do not dissolve in the solution. The precipitate formed here is lead iodide.
- If students did a pre-test at the beginning of the chapter, now is a good time to do the post-test. Be sure to immediately give students feedback on how they have improved.

Learning experience

What might happen when silver nitrate ($AgNO_3$) solution is added to potassium chloride (KCl) solution? Have the students predict what they think will happen. They should try to write the word equation and chemical equation using symbols. This reaction is a precipitation reaction where silver chloride (AgCl), the precipitant, forms a white solid. The chemical equation is:

$AgNO_3\,(aq) + KCl\,(aq) \rightarrow AgCl\,(s) + KNO_3\,(aq)$

As with the example given in this section, the reaction can be written more simply by showing only the ions which form the precipitate (the net ionic equation).

So:

$Ag^+\,(aq) + NO_3^-\,(aq) + K^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s) + K^+\,(aq) + NO_3^-\,(aq)$

becomes

$Ag^+\,(aq) + Cl^-\,(aq) \rightarrow AgCl\,(s)$

Learning experience

Students might like to find out more about precipitates. For example, they are often very colourful, and are often used as paint pigments. Precipitate reactions are used for removing salts from water in water treatment.

Learning experience

Creating a set of summary cards is an ideal activity for this chapter. On one side of each card, students put a diagram, question or unbalanced chemical equation; on the other side, they put a definition/explanation, answer or balanced equation. Get the students to shuffle the cards and place them on the table with the answer side facing down. Working in pairs, students choose a card and try to explain its meaning, answer the question or balance the equation. If they explain/answer it correctly, they keep the card. Students take turns in selecting a card. The winner is the person with the most correct cards.

Check! solutions

1 One way to solve the problem is to add a solution of potassium iodide to a small sample of each of the solutions and look for a yellow precipitate of lead iodide. Another alternative would be to allow a small sample to evaporate and look for any residue, which would indicate that a solute (lead nitrate) was present in the solution.

2 a The solubility of lead nitrate is 55 g/100 mL, and the solubility of silver chloride is 0.0002 g/100 mL.

b The most soluble compound is silver nitrate. The least soluble is silver iodide.

c The nitrates are the most soluble group.

d Magnesium forms the most soluble compounds.

e
- A precipitate of lead sulfate ($PbSO_4$) will form:
$Pb(NO_3)_2 + MgSO_4 \rightarrow PbSO_4 + Mg(NO_3)_2$
- A precipitate of silver iodide (AgI) will form:
$2AgNO_3 + MgI_2 \rightarrow 2AgI + Mg(NO_3)_2$
- A precipitate of magnesium carbonate ($MgCO_3$) will form:
$MgCl_2 + Na_2CO_3 \rightarrow MgCO_3 + 2NaCl$

f The best solutions to mix are silver nitrate and magnesium chloride.

3 The ionic equations are:

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

b $CuSO_4 \rightarrow Cu^{2+} + SO_4^{2-}$

c $Mg(NO_3)_2 \rightarrow Mg^{2+} + 2NO^{3-}$

d $Na_3PO_4 \rightarrow 3Na^+ + PO_4^{3-}$

Check!

1 Imagine you are agent 006 and have been captured by havoc agents. They have left you at a site in the desert formerly used to manufacture poisons. They have left two identical containers—one containing water and the other a colourless, odourless solution of toxic lead nitrate. You must drink soon or die of thirst. How could you find out which container holds the water? (There is a chemistry laboratory on the site.)

2 The table below gives the solubilities of various ionic compounds. The lower the solubility the more likely it is that the substance will form a precipitate. Silver chloride, which has a solubility of 0.0002 g/100 mL, is virtually insoluble and forms a precipitate, whereas magnesium nitrate, which has a solubility of 70 g/100 mL, is very soluble and would not form a precipitate.

a What is the solubility of:
- lead nitrate?
- magnesium carbonate?

Solubilities in g/100 mL water

	lead	magnesium	silver
carbonate	0.0002	0.01	0.002
chloride	1.0	55	0.0002
iodide	0.07	140	0.0000002
nitrate	55	70	220
sulfate	0.004	33	0.8

b Which of the compounds in the table is the most soluble in water? Which is the least soluble?

c Which group of compounds is the most soluble—the carbonates, chlorides, iodides, nitrates or sulfates?

d Which metal—lead, magnesium or silver—forms the most soluble compounds?

e Use the table to predict what will happen when you mix the following:
- lead nitrate and magnesium sulfate
- silver nitrate and magnesium iodide
- magnesium chloride and sodium carbonate.

Write a balanced equation for each reaction.

f Which compounds would you need to mix to form a precipitate of silver chloride?

3 Write ionic equations to show what happens when the following ionic compounds dissolve in water. (See page 255.)

a sodium chloride

b copper sulfate

c magnesium nitrate

d sodium phosphate

challenge

1 Lead nitrate and potassium iodide will not react unless they are dissolved in water. Write an inference to explain this.

2 In the reaction $PbI_2 \rightarrow Pb + I_2$ one element gains electrons and the other loses electrons. Explain how this happens.

3 Write a correctly balanced equation for each of the following reactions.

a Magnesium carbonate powder reacts with hydrochloric acid to produce a solution of magnesium chloride, carbon dioxide and water.

b Copper reacts with concentrated sulfuric acid (H_2SO_4) to produce a solution of copper(II) sulfate, sulfur dioxide and water.

c Lead(II) nitrate solution reacts with potassium chromate K_2CrO_4 to produce a yellow precipitate of lead chromate.

Challenge solutions

1 The substances lead nitrate and potassium iodide are ionic compounds that exist as solid crystals. The ions are held together in an arrangement called a lattice and will not react. However, if these compounds are dissolved in water, the ions are released, and are able to move about and react with each other to form a precipitate of lead iodide.

2 When lead iodide decomposes into its elements, the lead ions (Pb^{2+}) gain electrons to form lead atoms (Pb), and the iodide ions (I^-) lose electrons to form molecules of iodine (I_2).

3 These are the balanced equations:

a $MgCO_3 + 2HCl \rightarrow MgCl_2 + CO_2 + H_2O$

b $Cu + 2H_2SO_4 \rightarrow CuSO_4 + SO_2 + 2H_2O$

c $Pb(NO_3)_2 + K_2CrO_4 \rightarrow PbCrO_4 + 2KNO_3$

MAIN IDEAS

Copy and complete these statements to make a summary of this chapter. The missing words are on the right.

1 Atoms have no net charge. However, if they lose or gain ______ they become charged. They are then called ______.

2 Metals tend to ______ electrons to form positively charged ions. Non-metals tend to gain electrons to form ______ charged ions.

3 ______ bonds result from the attraction between ______ charged ions when metallic atoms transfer electrons to non-metallic atoms.

4 When an ionic compound ______ in water, the ions in the crystal lattice break apart and spread throughout the solution.

5 ______ bonds result from the sharing of electrons between non-metallic atoms.

6 The ______ or combining power of an atom tells you how it will combine with other atoms by losing, gaining or sharing electrons.

7 Chemical formulas can be used to write equations which represent reactions. These equations have to be ______ to give equal numbers of each type of atom on each side of the equation.

8 Atoms, molecules and ions are rearranged in chemical ______ to form new substances.

balanced
covalent
dissolves
electrons
ionic
ions
lose
negatively
oppositely
reactions
valency

Try doing the Chapter 10 crossword on the CD. Working with technology

REVIEW

1 What is the charge on an atom if it:
- a loses one electron?
- b gains two electrons?

2 a Which of the following are elements? How do you know?
- A C_2H_6O
- B NH_3
- C Cu
- D F^-
- E H_2
- F NO^{3-}

b Which represent compounds?
c Which represent ions?
d Which represent molecules?

3 What is the formula for a substance containing magnesium ions Mg^{2+} and hydroxide ions OH^-?
- A $MgOH_2$
- B $Mg(OH)_2$
- C MgOH
- D Mg_2OH

4 The formula for copper sulfate is $CuSO_4$.
- a How many different elements are there in copper sulfate?
- b What ions are formed when copper sulfate dissolves in water?
- c Is copper sulfate an ionic or a covalent compound?

Main ideas solutions

1 electrons, ions
2 lose, negatively
3 ionic, oppositely
4 dissolves
5 covalent
6 valency
7 balanced
8 reactions

Review solutions

1 a When an atom loses one electron its charge is 1+.
b When an atom gains two electrons its charge is 2–.

2 a C, D and E—since they contain only one type of atom
b A and B
c D and F
d A, B and E

3 B

2+ 1–

Mg_1 $(OH)_2$

$Mg(OH)_2$

4 a Three—Cu, S and O
b Cu^{2+} and SO_4^{2-} ions (see page 244)
c Ionic

5 What holds atoms together in:
  a an ionic compound?
  b a covalent compound?

6 Balance the following equations.
  a $C + Br_2 \rightarrow CBr_4$
  b $Fe_2O_3 + C \rightarrow Fe + CO$
  c $P_4 + H_2 \rightarrow PH_3$
  d $C_4H_8 + O_2 \rightarrow CO_2 + H_2O$
  e $Al_2(SO_4)_3 + Pb(NO_3)_2 \rightarrow PbSO_4 + Al(NO_3)_3$

7 Magnesium sulfate solution $MgSO_4$ is mixed with potassium hydroxide KOH.
  a What ions would be in the mixture?
  b Predict what new substances would be formed.
  c Write a balanced equation for the reaction you predict.

8 Why is it that copper sulfate solution conducts electricity but distilled water does not?

Fig 35 An iron nail in copper sulfate solution

9 The elements X, Y, Z and H form the following compounds: HX, YX_2 and YZ. Assuming H has a valency of 1+, what are the valencies of X, Y and Z?

10 a When nitric acid (HNO_3) is added to copper a reaction occurs and a blue solution is formed. Why is this so?
  b A brown gas called nitrogen dioxide is also formed. Infer whether the atoms in this gas come from the copper or the nitric acid. Explain your answer.

11 Kai placed an iron nail in some blue copper sulfate solution. The next day she noticed that the nail was partly dissolved and there were grains of copper on the bottom of the beaker. The solution had lost some of its blue colour. (See the photo below.)
  a Write an inference to explain why the solution lost some of its blue colour.
  b Use what you have learnt in Section 10.3 about ions changing partners to write a balanced equation for the reaction that occurred.

Check your answers on pages 337–338.

5 a Ionic compounds are held together by the attraction between oppositely charged ions (see page 237).
  b Covalent compounds are held together by the sharing of electrons (see page 240).

6 a $C + 2Br_2 \rightarrow CBr_4$
  b $Fe_2O_3 + 3C \rightarrow 2Fe + 3CO$
  c $P_4 + 6H_2 \rightarrow 4PH_3$
  d $C_4H_8 + 6O_2 \rightarrow 4CO_2 + 4H_2O$
  e $Al_2(SO_4)_3 + 3Pb(NO_3)_2 \rightarrow 3PbSO_4 + 2Al(NO_3)_3$

7 a $MgSO_4 \rightarrow Mg^{2+} + SO_4^{2-}$
  $KOH \rightarrow K^+ + OH^-$
  (See the table on page 172.)
  b magnesium sulfate + potassium hydroxide →
  magnesium hydroxide + potassium sulfate
  c $MgSO_4 + 2KOH \rightarrow Mg(OH)_2 + K_2SO_4$

8 When copper sulfate dissolves in water, Cu^{2+} and SO_4^{2-} ions are formed. These ions are free to move and conduct an electric current.
  Distilled water contains uncharged H_2O molecules which do not conduct an electric current.

9 X has a valency of 1–. (You can tell this from the formula HX, since H has a valency of 1+.)
  Y has a valency of 2+ (because of YX_2)
  Z has a valency of 2– (because of YZ)

10 a The nitric acid reacts with the copper to form copper ions which make the solution blue.
  b Nitrogen dioxide (formula NO_2) contains nitrogen and oxygen atoms. These atoms must have come from the nitric acid (HNO_3).

11 a The blue colour is due to copper ions in solution. Since copper was produced in the reaction, you can infer that the copper ions changed to copper atoms. This is why the solution lost some of its blue colour.
  b If copper is formed then you are left with iron sulfate.
  copper sulfate + iron → copper + iron sulfate
  $CuSO_4 + Fe \rightarrow Cu + FeSO_4$
  (assuming a valency of 2+ for iron)
  or $3CuSO_4 + 2Fe \rightarrow 3Cu + Fe_2(SO_4)_3$
  (assuming a valency of 3+ for iron)