Experiment worksheet

3.3 Groups in the periodic table have properties in common

Pages 74–75 and 195

Experiment 3.3: Reactivity of metals

Aim

To compare the reactivity of various metals by observing their reaction with hydrochloric acid.

Materials

• 2 M hydrochloric acid

• Detergent

• 0.5 cm pieces of magnesium, aluminium, iron, zinc and copper

• Steel wool

• Test tubes and test-tube rack

• Ruler

• Timer

• Bench mat



CAUTION: Wear protective gloves and safety glasses throughout this experiment.

Method

1 Clean the surface of the magnesium with a piece of steel wool.

2 Place the magnesium into a test tube.

3 Add 3 drops of detergent to the test tube.

4 Add 2 cm of hydrochloric acid to the test tube. Set the timer for 5 minutes and record your observations, including the height of the foam produced, in a results table like the one below.

5 Repeat the process for the remaining metals.

6 Record your observations over 30 minutes.

7 This equipment can be left set up overnight to observe any further changes.

Results

Copy and complete the following table.

|  |  |  |
| --- | --- | --- |
| Metal | Observations | Height of foam (cm) |
| Magnesium |  |  |
| Aluminium |  |  |
| Iron |  |  |
| Zinc |  |  |
| Copper |  |  |

Discussion

1 Which metal was the most reactive?

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2 Which metal was the least reactive?

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3 Why were the metals cleaned with steel wool first?

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4 Why was detergent added to the test tubes with the hydrochloric acid?

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5 What properties would you think the most reactive metal would also exhibit?

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6 Is there any link between the reactivity of the metals and where they are located in the periodic table?

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Conclusion

What do you know about the reactivity of metals?

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Experiment worksheet

3.4 Non-metals have properties in common

Pages 76–77 and 196

Challenge 3.4: Identifying patterns in the periodic table

What you need

• A3 sheet of paper

• Pens

• Highlighter pens

What to do

1 On an A3 sheet of paper, make a copy of the periodic table up to element 20. Leave a gap for the block of transition metals. Ensure that the size of the box for each element will fit the information you will need to insert, as detailed below. (Chlorine has been completed in step 5 as an example.)

2 Colour hydrogen red, metals blue, noble gases purple, other non-metals green, and use the different colours to shade in the metals, the noble gases, hydrogen, the non-metals other than the noble gases and hydrogen, and the metalloids. Place a suitable key under your periodic table.

3 Identify the elements that will not gain or lose electrons in a reaction because their uncharged atoms are already very stable. Beneath them, write:

• ‘Already a stable structure’.

• ‘Does not form an ion’.

4 Identify the elements that will not gain or lose electrons in a reaction, because this would require them to gain or lose more than three electrons. Beneath them, write:

• ‘Needs to gain or lose more than three electrons for a more stable structure’

• ‘Does not form an ion’.

5 Complete the box for each of the other elements listed, except for the metalloids and hydrogen, by stating how many electrons the element needs to gain or lose to achieve a more stable structure, and hence what charge its ion should have, like the example of chlorine. Information for chlorine:

• Chlorine (Cl)

• Needs to gain one electron

• Charge on ion = –1

Discussion

1 What patterns do you notice in your entries for the alkali metals?

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2 What patterns do you notice in your entries for the alkaline earth metals?

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3 What patterns apply to all the metals listed?

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4 What patterns do you notice in your entries for the halogens?

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5 What patterns do you notice in your entries for the group 16 elements?

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| --- |
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6 What patterns apply to the non-metals, except for hydrogen and the noble gases?

7 In general, what do you expect to happen when a metal atom and a non-metal atom meet? Which groups of non-metals will not react in this way? Discuss.

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8 Predict what might happen if a:

a potassium atom and a fluorine atom meet

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|  |

b calcium atom and an oxygen atom meet.

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9 You can illustrate your predictions by drawing shell diagrams of the atoms and showing what happens in the reaction.

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10 Suggest why hydrogen and the metalloids were not considered in this activity.

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Experiment worksheet

3.5 Metal cations and non-metal anions combine to form ionic compounds

Pages 78–79 and 197

Experiment 3.5: Conductivity of ionic compounds

Aim

To investigate the electrical conductivity of two ionic compounds as a solid and in aqueous solution.

Materials

• Large sodium chloride crystals

• Coarse sea salt crystals

• Small Petri dish

• 4 V battery or other 4 V DC power source

• Ammeter

• Wires with alligator clips

• 2 graphite electrodes

• 3 × 100 mL beakers

• Large spatula

• Glass stirring rod

• Paper towel

Method

1 Set up the electrical circuit as shown in Figure 1. Have your teacher check that it is correct before proceeding. Ensure that you know how to use the ammeter and its scales correctly.

****  
Figure 1 Experiment set-up.

2 Using the spatula, place the largest sodium chloride crystal onto the Petri dish, then touch each end with an electrode, making sure that the two electrodes do not touch each other. Does the crystal conduct electricity? If it doesn’t appear to, connect the wire to the more sensitive scale on the ammeter. Does a reading register now? Record your result.

3 In a 100 mL beaker, place half a large spatula of sodium chloride crystals and add 50 mL of water. Stir to dissolve the crystals.

4 Place the electrodes into this solution, again ensuring they do not touch each other. Does the solution conduct electricity? If it doesn’t appear to, connect the wire to the more sensitive scale on the ammeter. Does a reading register now? Record your result.

5 Turn off the power supply and rinse the electrodes with fresh tap water, then dry them with a paper towel.

Inquiry

What if large coarse sea salt was used?

1 Write a hypothesis for your question.

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2 What is your independent variable?

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3 What is your dependent variable?

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4 List three variables that you will need to control. How will you control them?

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5 Record your observations in a table.

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Results

Devise a simple table or spreadsheet in which to record your results.

Discussion

1 Sea salt is a mixture of different ionic compounds, including sodium chloride. What can you conclude about the ability of solid ionic compounds to conduct electricity, whether they are pure or mixed together?

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2 What effect does dissolving an ionic compound in water have on its ability to conduct electricity?

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3 To conduct electricity, a substance must have charged particles that can move about. Suggest an explanation for your findings.

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4 The melting point of sodium chloride is 801°C, so it is not practical to melt it in the school laboratory. Predict whether molten sodium chloride would conduct electricity and justify your answer.

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Conclusion

5 What do you know about the conductivity of ionic compounds?

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Experiment worksheet

3.5 Metal cations and non-metal anions combine to form ionic compounds

Pages 78–79 and 198

Skills lab 3.5: Ionic compounds

Ionic compounds are those formed from the bonding of ions. Consider sodium chloride, which is produced when sodium and chlorine meet and react. In this compound, the metal sodium is present in the form of positively charged ions (Na+) and the non-metal chlorine is present as negatively charged ions (Cl–).  
Notice that the:

• metal is named first and its name is not changed

• non-metal is named second and the end of its name is changed from -ine to -ide.

This obeys the following standard naming convention.

• The positively charged ion (cation) in the compound is written first and keeps the name of the metal from which it was formed.

• The negatively charged ion (anion) in the compound is written second. The end of the name of the non-metal from which it formed is replaced with -ide.

• Some transition metals can form more than one ion. In these cases, a Roman numeral is used to show the charge on the ion. For example, copper forms two ions: one with a 1+ charge and one with a 2+ charge. These ions are called copper(I) and copper(II) ions, respectively.

|  |  |  |  |
| --- | --- | --- | --- |
| Cations | | Anions | |
| Name | Formula | Name | Formula |
| Lithium | Li+ | Fluoride | F− |
| Sodium | Na+ | Chloride | Cl− |
| Potassium | K+ | Bromide | Br− |
| Magnesium | Mg2+ | Iodide | I− |
| Calcium | Ca2+ | Oxide | O2− |
| Aluminium | Al3+ | Sulfide | S2− |
| Silver | Ag+ | Nitride | N3− |
| Zinc | Zn2+ |  |  |
| Copper(II) | Cu2+ |  |  |
| Iron(II) | Fe2+ |  |  |
| Iron(III) | Fe3+ |  |  |

The formula for sodium chloride is NaCl, whereas the formula of magnesium chloride is MgCl2. The formula NaCl means that the cations and anions are present in a ratio of 1:1. That is, for every Na+ ion present in a sodium chloride crystal, there is one Cl– ion present. The formula MgCl2 means that the cations and anions are present in a ratio of 1:2. That is, for every Mg2+ ion present in a magnesium chloride crystal, there are two Cl– ions present. This is necessary to achieve an overall neutral charge.

We can use this principle to determine the formula of an ionic compound. First, use Table 8.1 to list the formulas of the cations and anions present. Then, work out the simplest ratio they need to be in so that the total positive charge and total negative charge are equal.

Example

1 What is the formula for iron(II) oxide?

• The ions are Fe2+ and O2–.

• Because the charges 2+ and 2– are equal, the ions only need to be in a ratio of 1:1.

• Therefore, the formula is FeO.

2 What is the formula for silver sulfide?

• The ions are Ag+ and S2–.

• Because the charges are 1+ and 2–, the ions need to be in a ratio of 2:1 (making it a total of 2+ and 2–).

• Therefore, the formula is Ag2S.

Your turn

Write the formulas for:

|  |  |
| --- | --- |
| a lithium bromide |  |
| b iron(III) chloride |  |
| c sodium nitride |  |
| d aluminium oxide |  |

Experiment worksheet

3.6 Non-metals combine to form covalent compounds

Pages 80–81 and 199

Challenge 3.6: Modelling covalent molecules

Aim

To model the sharing of electrons in covalent molecules.

What you need

• Molecular modelling kits (or use different coloured marshmallows and toothpicks)

What to do

1 Choose three different colours to represent carbon, hydrogen and oxygen.

2 For each of the molecules shown in the results table:

a state the numbers of each atom

b make and draw a model of the molecules

c draw the number of electrons in the valency shell of each atom including the shared electrons.

Results

Copy and complete the following table.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Molecule | Atoms present | Numbers of each atom | Drawing of model | Electron dot diagrams |
| H2 |  |  |  |  |
| H2O |  |  |  |  |
| CH4 |  |  |  |  |
| CO2 |  |  |  |  |
| CHCl3 |  |  |  |  |

Discussion

1 What type of bond occurs between a metal and a non-metal?

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2 What type of bond occurs between two non-metals?

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3 What is a valency shell?

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4 What is meant by the term ‘sharing electrons’ in covalent bonds?

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Conclusion

What do you know about covalent bonds?

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Experiment worksheet

3.7 Metals form unique bonds

Pages 82–83 and 199

Challenge 3.7: Modelling alloys

Aim

To compare the properties of model alloys.

What you need

• 4 different colours of plasticine (35 g)

• Sand (12 g)

• Newspaper

• Balance

• Magnifying glass

What to do

1 Weigh 2 g of sand onto the newspaper.

2 Roll and work one of the plasticine colours until it is soft and malleable. Roll it out into a 0.5 cm layer.

3 Sprinkle the sand onto the plasticine and roll it over the sand until the sand it evenly spread through.

4 Repeat steps 2 and 3 with 4 g and 6 g of sand.

5 Work and shape the four pieces of plasticine until they are at room temperature.

6 Form each shape into a cylinder of the same size and length.

7 Hold the ends of a plasticine cylinder firmly and pull firmly apart.

8 Repeat the pull test for each plasticine cylinder.

9 Use the magnifying glass to examine the broken ends of the cylinder.

Results

Record your observations in an appropriate table.

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Discussion

1 Which ‘alloy’ was most malleable (able to be rolled out easily when cold)?

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| --- |
|  |

2 Which ‘alloy’ was most ductile (able to be drawn out into a wire easily)?

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3 Which ‘alloy’ was most brittle (snapped quickly)?

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4 Did the amount of sand in the ‘alloy’ affect the size of the largest fracture surface? Explain your observation.

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Conclusion

How does the alloying of metal affect its properties?

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