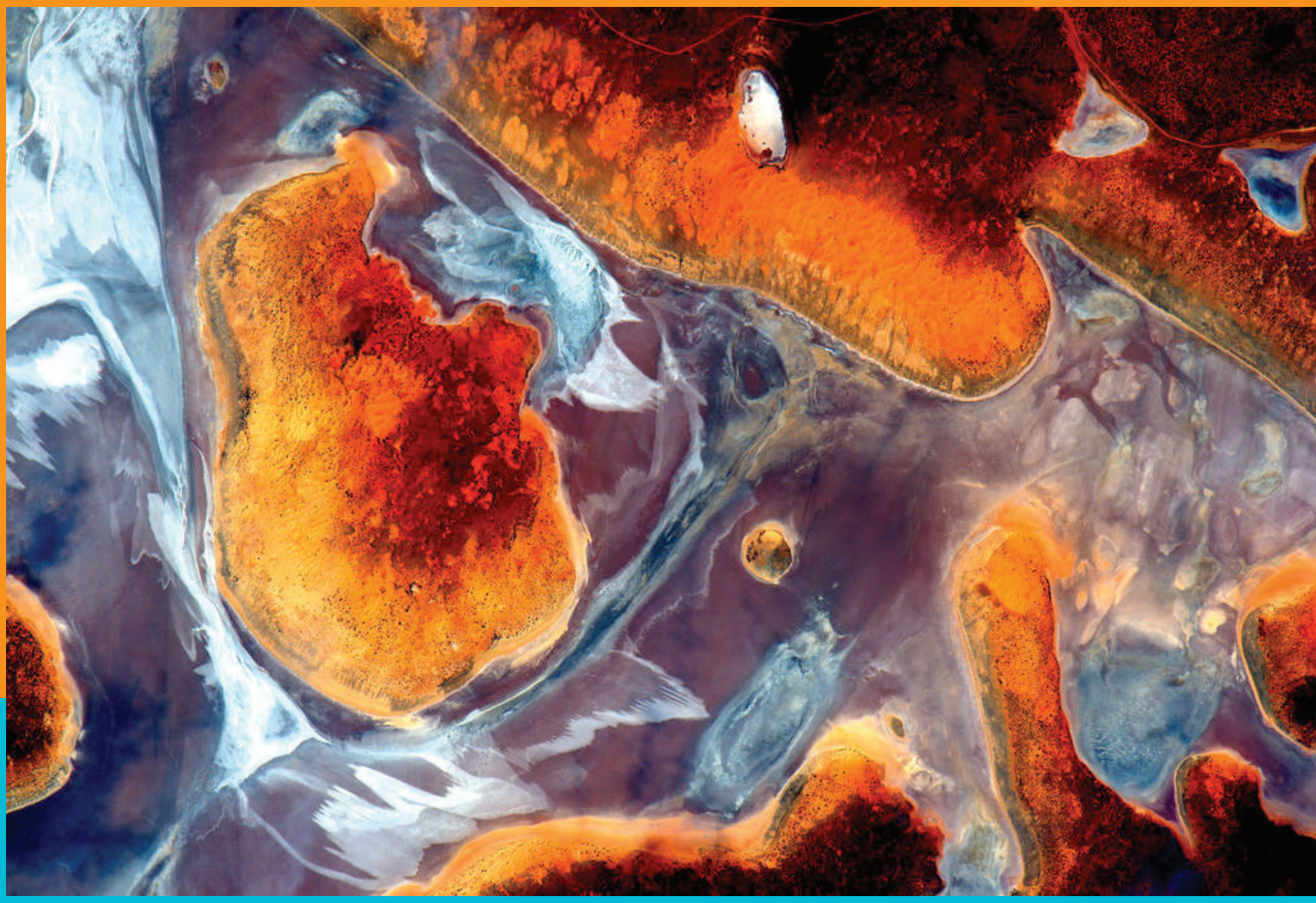

EDEXCEL INTERNATIONAL GCSE (9–1)

CHEMISTRY

TEACHER RESOURCE PACK

Chapter 7: Ionic compounds

Ions and ionic bonding

1. Use the words in the box to complete the gaps below. Use each word only once.

Missing words: anion, cation, charge, decreases, electrons, ion, ionic, metal, negative, non-metals, oppositely, outer, positive

Ionic compounds usually contain a _____ and a non-metal. An _____is an element or compound that has lost or gained _____. Metals lose electrons to form _____ ions. Non-metals form negative ions by _____ electrons. The charge of the metals and non-metals change because electrons have a _____ charge. If electrons are gained the overall negative _____increases. If electrons are lost the negative overall charge_____. A positively ion is known as a _____ and a negative ion is known as an _____. To form an _____ bond, electrons move from a metal atom to the non-metal atom. This produces a positive and a negative ion. The _____ charged ions then attract each other. Electrons are transferred from the metals to the _____ to complete the non-metals atom's _____ electron shell.

The structure of ionic compounds

2. Draw dot and cross diagrams for the following compounds:
   a. Aluminium oxide (Al_2O_3)

   b. Iron chloride ($FeCl_2$)

   c. Potassium nitrate (KNO_3)

3. Add the symbols for the ions formed by the following elements and compounds.

Element/compound	Ion formed	Element/compound	Ion formed
Calcium		Iodine	
Barium		Silver	
Aluminium		Ammonium	
Beryllium		Sulfate	
Phosphorous		Nitrate	

Ionic compounds

4. Give the formula for the following ionic compounds.
   a. Aluminium hydroxide
   b. Barium sulfate
   c. Ammonium chloride
   d. Calcium carbonate
   e. Copper(II) chloride
   f. Lead(II) nitrate

Properties of ionic compounds

5. Look at the table below. Use the data to decide which of the compounds are ionic and which are not.

Compound	Melting point /^{0}C	Boiling point /^{0}C	Electrical conductivity when molten	Electrical conductivity when in aqueous solution
P	2015	2980	Good	Good
Q	-87	-67	Poor	Poor
R	300	500	Good	Insoluble
S	558	1506	Good	Good

The ionic compound(s) are: _____

Chapter 5: Chemical formulae, equations and calculations Part 1

Alignment with Student Book: Pages 38-63

Chapter overview

This chapter introduces quantitative chemistry. Students will explore chemical formulae, including balancing equations. There will be a focus on the use of calculations including relative molecular mass and relative atomic mass. The use of experimental data will feature prominently and students will use a variety of techniques including reacting masses and the determination of formula by combustion. The different types of formula will be introduced for the first time.

What to expect

1.25 write word equations and balanced chemical equations (including state symbols):
• for reactions studied in this specification
• for unfamiliar reactions where suitable information is provided

1.26 calculate relative formula masses (including relative molecular masses) (Mr) from relative atomic masses (Ar)

1.27 know that the mole (mol) is the unit for the amount of a substance

1.28 understand how to carry out calculations involving amount of substance, relative atomic mass (Ar) and relative formula mass (Mr)

1.29 calculate reacting masses using experimental data and chemical equations

1.30 calculate percentage yield

1.31 understand how the formulae of simple compounds can be obtained experimentally, including metal oxides, water and salts containing water of crystallisation

1.32 know what is meant by the terms empirical formula and molecular formula

1.33 calculate empirical and molecular formulae from experimental data

1.36 practical: know how to determine the formula of a metal oxide by combustion (e.g. magnesium oxide) or by reduction (e.g. copper(II) oxide)

This chapter contains material that is much more complex. Due to the quantitative nature of much of the content students will need to have relatively strong mathematical skills.

Balancing equations should not be new to students and so will not need much time. However, teaching how to calculate relative atomic mass and relative formula mass will need ample time set aside. The mole has been a regularly identified topic which students either do not fully understand or find very boring. Though the use of the mole in this chapter is limited, it cannot be emphasised strongly enough how important it is for students to fully comprehend what the mole is and how it is used. As much practice should be given as possible, either in class or for homework, using the different equations and calculations.

As there is so much practical work, lesson time must be considered so as not to rush the calculation aspect of the investigation.

Teaching notes

Starter Activities

Elements and compounds states game - Students must write three lists. One for each of the three states of matter and list as many elements or compounds they can name that occur in each state at room temperature.

The mole demonstration - Weigh out 1 mole of a variety of different elements to reinforce that although the substances have different masses they have the same number of particles inside (Carbon 12g, water 16g, Magnesium 24g, copper(II) carbonate 124g). Ask students if they recognise the numbers for the elements from anywhere? Are they on the periodic table? The mass of 1 mole is equal to the relative atomic mass.

Main activities/practical work

Balancing equations practice - Give students a variety of unbalanced equations. This can be differentiated very easily with some students given more complex examples.

The change in mass when magnesium burns demonstration or practical - Students weigh some magnesium ribbon then burn it in air. They reweigh the new compound formed and use the result to determine the formula of magnesium oxide. See activity three on page 48 of the textbook.

Determining relative atomic mass practical - Students measure the volume of hydrogen gas produced when magnesium ribbon reacts with hydrochloric acid to determine the relative atomic mass of magnesium.

Finding the formula of hydrated copper(II) sulfate practical - Students weigh some hydrated copper(II) sulfate and then heat it to remove the water. They reweigh the copper(II) sulfate to find the water of crystallisation. Mole calculations are then used to find the formula. See page 53 of the textbook.

Finding the formula of copper oxide using methane practical or demonstration - Copper(II) oxide is reduced using methane gas which is passed over the oxide as it is heated. This is a complicated practical and so may be better as a demonstration with some students. By weighing the Copper(II) oxide and then the copper produced the formula of copper(II) oxide can be calculated. See activity 4 on page 49 of the textbook.

Determining the formula of water demonstration- See page 51 of the textbook.

Homework

The questions in the book on page 60-63 provide enough challenge to be set for a number of homework sessions. The homework sheet in the TRP offers additional questions.

Possible misunderstandings

As students may struggle with the concept of the mole, try to use the analogy of a dozen. A dozen apples and a dozen bananas do not have the same mass but there is the same number of each fruit. Students may confuse the different formulae they need to use so care must be taken to ensure they know when to use each.

When completing practical work remind students to weigh their reactants and their products as they often forget to weigh the reactants which then prevents any kind of analysis.

Highlight anything that students might find difficult to understand in the chapter in more detail. Provide clarity on the issues. Suggest ways in which teachers can explain or demonstrate the content so that it is most clear.

Differentiation

Students may be given a fourth list for the start game to include substances in solution (aq).

There is extension work on page 46 of the textbook on the Avogadro constant.

The use of formula triangles can be used to help support students. It is very easy for students to get lost as they progress through worked examples. Frequent learning checks are necessary.

Practicals

The change in mass when magnesium burns demonstration or practical. Details may be found here:

http://www.rsc.org/learn-chemistry/resource/res00000718/the-change-in-mass-when-magnesium-burns

Determining relative atomic mass practical - Details may be found here:

http://www.rsc.org/learn-chemistry/resource/res00000401/determination-of-relative-atomic-mass?cmpid=CMP00006706

Finding the formula of hydrated copper(II) sulfate practical- Details may be found here:

http://www.rsc.org/learn-chemistry/resource/res00000436/finding-the-formula-of-hydrated-copper-ii-sulfate

Finding the formula of copper oxide using methane practical or demonstration- Details may be found here:

http://www.rsc.org/learn-chemistry/resource/res00000727/finding-the-formula-of-copper-ii-oxide

Unit 2 Multiple-choice questions

1. Look at the table below. Which one of the unknown elements is most likely to be potassium?

Substance	Reaction with air	Reaction with water
A	No reaction	No reaction
B	Tarnishes slowly	Slow effervescence
C	Tarnishes quickly	Effervescence
D	Tarnishes immediately	Violent effervescence

2. Which of the following statements best explains why group 1 metals become more reactive down the group?

   A. Going down the group the distance between the nucleus and the outer shell electron decreases. This means that the force of attraction is weaker and therefore the electron is more easily lost.
   B. Going down the group the distance between the nucleus and the outer shell electron increases. This means that the force of attraction is stronger and therefore the electron is more easily lost.
   C. Going down the group the distance between the nucleus and the outer shell electron increases. This means that the force of attraction is weaker and therefore the electron is more easily lost.
   D. Going down the group the distance between the nucleus and the outer shell electron increases. This means that the force of attraction is stronger and therefore the electron is less easily lost.

3. Where in the Periodic Table would the halogens be found?

   A. Group 2
   B. Group 5
   C. Group 7
   D. Group 8

4. Halogens vary in their reactivity. Based on your knowledge which of the following reactions would not take place?

   A. Potassium bromide + Chlorine -> Potassium chloride + Bromine
   B. Sodium chloride + Iodine -> Sodium iodide + Chlorine
   C. Magnesium Iodide + Bromine -> Magnesium bromide + Iodine
   D. Zinc Bromide + Chlorine -> Zinc chloride + Bromine

5. Chlorine is one of the halogens. Look at the descriptions below and select the one the best describes chlorine at room temperature.

   A. Yellow gas

B. Grey solid
C. Red/brown liquid
D. Green gas

6. Astatine is a halogen and can be found in the Periodic Table, underneath Iodine. Using your knowledge of the trends of the halogens choose the most suitable description of astatine's properties.

A. Very reactive gas.
B. Very reactive liquid.
C. Unreactive solid.
D. Unreactive liquid.

7. Why is chlorine more reactive than iodine?

A. Chlorine has a stronger tendency for form a 1- ion as its nucleus is closer to its outer electron shell.
B. Iodine has a stronger tendency for form a 1- ion as its nucleus is further from its outer electron shell.
C. Chlorine has a weaker tendency for form a 1- ion as its nucleus is closer to its outer electron shell.
D. Iodine has a stronger tendency for form a 1- ion as its nucleus is closer to its outer electron shell.

8. The percentage by volume of oxygen in the atmosphere is?

A. 78%
B. 0.4%
C. 21%
D. 0.9%

9. Which of the following would be a suitable test for oxygen?

A. Turns limewater cloudy.
B. Relights a glowing splint.
C. Makes a squeaky pop when ignited.
D. Turns damp litmus paper blue.

10. Burning fossil fuels can release oxides into the atmosphere. Which one of the following oxides is not acidic?

A. Carbon dioxide.
B. Nitrogen dioxide.
C. Sulfur dioxide.
D. Potassium oxide.

11. When heated, copper (II) carbonate thermally decomposes. Which of the following reactions shows the correct products?

    A. Copper (II) carbonate -> Copper hydroxide + Carbon
    B. Copper (II) carbonate -> Copper oxide + Carbon dioxide
    C. Copper (II) carbonate -> Copper oxide + Hydrogen
    D. Copper (II) carbonate -> Copper hydroxide + Carbon dioxide

12. Magnesium burns in oxygen to form magnesium oxide. Choose the correct symbol equation for this reaction.

    A. $2Mg + O_2 \rightarrow 2MgO$
    B. $Mg + O_2 \rightarrow MgO$
    C. $Mg + O \rightarrow MgO$
    D. $2Mg + O_2 \rightarrow MgO_2$

13. Which of the following techniques could not be used to determine the percentage of oxygen in air?

    A. By reacting wet iron filings inside a conical flask and measuring the volume decrease using a gas syringe.
    B. By burning wire wool on a balance to show the mass increase.
    C. By passing a known volume of air back and forth across heated copper turnings inside a silica tube.
    D. By reacting phosphorus inside a bell jar set in water and measuring the change in water level.

14. Which of the following statements about reduction is true?

    A. Reduction is the gain of electrons or the loss of hydrogen.
    B. Reduction is the gain of electrons or the loss of oxygen.
    C. Reduction is the loss of electrons or the loss of oxygen.
    D. Reduction is the gain of electrons or the gain of oxygen.

15. A piece of magnesium is added to copper sulfate solution. Which of the following does not happen?

    A. The magnesium becomes covered in a thin layer of orange copper.
    B. The colour of the solution becomes less blue.
    C. Magnesium sulfate is formed.
    D. There is a drop in the temperature of the solution.

16. Choose the conditions under which iron will rust at the fastest rate.

    A. Dry, warm and in the presence of salt.
    B. Dry, cold and in the presence of salt.

C. Wet, warm and in the presence of salt.
D. Wet, warm and in the absence of salt.

17. The reactivity series is used to show the difference in the reactivity of elements. Choose the answer showing elements in the correct order of reactive from most reactive to least reactive.

Most reactive Least reactive

A. Gold – sodium – magnesium- copper – zinc - iron
B. Potassium – zinc – sodium- calcium – aluminium - copper
C. Sodium - magnesium- aluminium - iron – copper – gold
D. Lithium – aluminium- zinc- iron – copper – silver

18. Bocks of magnesium can be added to the hull of ships made from iron to reduce the amount of rusting that takes place. Which type of protection is this?

A. Sacrificial protection.
B. Barrier protection.
C. Galvanising.
D. Vulcanising.

19. Choose the correct statement about reducing agents.

A. Reducing agents gain electrons and are therefore oxidised during reactions.
B. Reducing agents gain electrons and are therefore oxidised during reactions.
C. Reducing agents give away electrons and are therefore reduced during reactions.
D. Reducing agents give away electrons and are therefore oxidised during reactions.

20. When added to dilute acids metals react in a general way. Which of the following shows the correct products of this type of reaction?

A. Metal + Acid -> Salt + Water
B. Metal + Acid -> Salt + Hydrogen
C. Metal + Acid -> Metal oxide + Water
D. Metal + Acid -> Salt + Water + Carbon dioxide

21. Metals are usually found combined with other elements in rocks called ores. A few unreactive metals are found in their pure form. Choose the metal that is most likely to be found in its pure form.

A. Aluminium.
B. Iron.
C. Silver.
D. Lead.

22. Why is carbon used in the extraction of iron form iron oxide?

    A. Because it is cheap.
    B. Because it is more reactive than iron.
    C. Because it is less reactive than iron.
    D. Because it is more reactive than oxygen.

23. Copper and its alloys are widely used metals. The use of copper often depends on a specific property it has. Pick the correct use of copper with the property that makes it suitable.

    A. Wires- Copper is a good conductor of heat.
    B. Water pipes- Copper is a very reactive metal.
    C. Surfaces in hospitals- Copper has antimicrobial properties.
    D. Pots and pans- Copper is a good conductor of electricity.

24. Steel is an example of an alloy. Select the best definition of an alloy.

    A. A mixture of a metal with, usually, other metals or carbon.
    B. A compound of a metal with a non-metal.
    C. A metal produced by electrolysis.
    D. A mixture of a metal with carbon.

25. Which of the following metals could not be extracted from its oxide by reduction with carbon?

    A. Copper.
    B. Iron.
    C. Zinc.
    D. Aluminium.

26. Mild steel is an alloy of iron which contains about 0.25% carbon. Which of the following is not a common use of mild steel?

    A. Cutlery.
    B. Car bodies.
    C. Nails.
    D. Bridges.

27. Alloys are often used because they are harder than pure metals. Why are alloys harder?

    A. In alloys, the different sized atoms in the lattice make it harder for the layers of ions to slide over one another.
    B. In alloys, the different sized atoms in the lattice make it easier for the layers of ions to slide over one another.
    C. In alloys, the different sized electrons in the lattice make it harder for the layers of ions to slide over one another.

D. In alloys, the atoms are the same size which makes it harder for the layers of ions to slide over one another.

28. An indicator is a substance that can be used to determine the pH of an acid or alkali. A number of indicators were added to a sample of acid. Choose the row from the table that shows the results you would expect to see for each indicator if it was added to an acid.

Sample	Phenolphthalein	Methyl orange	Universal indicator	Litmus
A	Colourless	Red	Red	Red
B	Red	Red	Red	Red
C	Pink	Red	Blue	Blue
D	Pink	Orange	Red	Blue

29. Ethanoic acid is a weak acid. What pH value would you expect it to have?

A. 12
B. 4
C. 2
D. 7

30. A wasp sting is alkaline and can be neutralised using a weak acid like vinegar. Which of the following description about acids and alkalis is correct?

A. In an aqueous solution, alkalis are sources of hydrogen ions and acids are sources of hydroxide ions.
B. In an aqueous solution, acids are sources of hydrogen ions and alkalis are sources of hydroxide ions.
C. In an aqueous solution, acids are sources of hydroxide ions and alkalis are sources of hydrogen ions.
D. In an aqueous solution, acids are sources of hydrogen ions and alkalis are sources of hydrogen ions.

31. Two students were testing the solubility in water of different nitrates. Which row shows the results that you would expect to obtain?

	Potassium nitrate	Sodium nitrate	Calcium nitrate
A	Soluble	Insoluble	Soluble
B	Insoluble	Insoluble	Soluble
C	Insoluble	Soluble	Insoluble
D	Soluble	Soluble	Soluble

32. Which of the following chlorides are insoluble in water?

A. Silver chloride.
B. Calcium chloride.

C. Sodium chloride.
D. Copper chloride.

33. Calcium carbonate is a base that is often used in medication to reduce stomach acidity. The acid present in your stomach is hydrochloric acid. Choose the correct balanced equation for the reaction between calcium carbonate and hydrochloric acid.

A. $CaCO_3 + HCl \rightarrow CaCl_2 + CO + H_2O$
B. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2$
C. $CaCO_3 + HCl \rightarrow CaCl_3 + CO_2 + H_2O$
D. $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$

34. Copper (II) sulfate crystals can be made from copper oxide. Choose the best method from the answers below.

A. Add copper oxide to hot hydrochloric acid until no more will dissolve, filter off the undissolved copper oxide, heat the filtrate in an evaporating basin until blue crystal begin to form.
B. Add copper oxide to hot sulfuric acid until no more will dissolve, filter off the undissolved copper oxide, distil the filtrate into a condensing tube.
C. Add copper oxide to hot sulfuric acid until no more will dissolve, filter off the undissolved copper oxide, heat the filtrate in an evaporating basin until blue crystal begin to form.
D. Heat the copper oxide with powdered carbon, dissolve both in water and then filter off any undissolved copper oxide, leave the filtrate to cool in an evaporating basin.

35. Which of the following is the correct colour for a lithium flame test?

A. Blue.
B. Pink.
C. Red.
D. Lilac.

36. Ammonia is a strong-smelling gas which can be harmful if inhaled in high concentrations. What is the chemical test for ammonia?

A. Heat the liquid and hold a piece of damp litmus paper at the end of the test tube, it will turn blue.
B. By bubbling it through lime water, it will turn cloudy.
C. By adding a glowing splint, it will reignite.
D. Heat the liquid and hold a piece of litmus paper at the end of the test tube, it will be bleached.

37. What would you expect to see if copper (II) sulfate was added to sodium hydroxide solution?

A. A yellow precipitate forms.
B. Effervescence.

C. A blue precipitate.
D. A white precipitate.

38. Look at the table below. Which row shows the correct results when iron (II) and iron (III) ions are added to sodium hydroxide?

Sample	Iron (II) ions	Iron (III) ions
A	Orange/ brown precipitate	Orange/ brown precipitate
B	Green precipitate	Orange/ brown precipitate
C	Green precipitate	Green precipitate
D	Orange/ brown precipitate	Green precipitate

39. Flame tests are a useful way to identify unknown substances. A sample of an unknown powder was added to a damp splint. The split was then held in a flame. A yellow flame was observed. Which ion was present?

A. Lithium
B. Calcium
C. Copper (II)
D. Sodium

40. Describe the colour change observed when water is added to anhydrous copper (II) sulfate.

A. Blue to white.
B. Blue to red.
C. White to blue.
D. White to orange.

Answers

1. D
2. C - Going down the group the distance between the nucleus and the outer shell electron increases. This means that the force of attraction is weaker and therefore the electron is more easily lost.
3. C - Group 7
4. B - Sodium chloride + Iodine -> Sodium iodide + Chlorine
5. D - Green gas
6. C - Unreactive solid.
7. A - Chlorine has a stronger tendency for form a 1- ion as its nucleus is closer to its outer electron shell.
8. C - 21%
9. B - Relights a glowing splint.
10. C - Potassium oxide.
11. B - Copper (II) carbonate -> Copper oxide + Carbon dioxide
12. A - $2Mg + O_2 \rightarrow 2MgO$
13. B - By burning wire wool on a balance to show the mass increase.
14. B - Reduction is the gain of electrons or the loss of oxygen.
15. D - There is a drop in the temperature of the solution.
16. C - Wet, warm and in the presence of salt.
17. C - Sodium - magnesium- aluminium - iron – copper – gold
18. A - Sacrificial protection
19. D - Reducing agents give away electrons and are therefore oxidised during reactions.
20. B - Metal + Acid -> Salt + Hydrogen
21. C - Silver
22. B - Because it is more reactive than iron.
23. C - Surfaces in hospitals- Copper has antimicrobial properties.
24. A - A mixture of a metal with, usually, other metals or carbon.
25. D - Aluminium
26. A - Cutlery
27. A - In alloys, the different sized atoms in the lattice make it harder for the layers of ions to slide over one another.
28. A
29. B - 4
30. B - In an aqueous solution, acids are sources of hydrogen ions and alkalis are sources of hydroxide ions.
31. D
32. A - Silver chloride.
33. D - $CaCO_3 + 2HCl \rightarrow CaCl_2 + CO_2 + H_2O$
34. A - Add copper oxide to hot hydrochloric acid until no more will dissolve, filter off the undissolved copper oxide, heat the filtrate in an evaporating basin until blue crystal begin to form.
35. C - Red
36. A - Heat the liquid and hold a piece of damp litmus paper at the end of the test tube, it will turn blue.
37. C - A blue precipitate.
38. B
39. D - Sodium
40. C - White to blue.

Chemistry paper 1
Exam question

- Copper oxide can be used to make copper(II) sulfate using the reaction below.
- $CuO(s) + H_2SO_4(aq) \rightarrow CuSO_4(aq) + H_2O(l)$

- Explain how a student could use this reaction to obtain a sample of pure hydrated copper(II) sulfate crystals if supplied with copper(II) oxide and dilute sulfuric acid.

(5)

Student response 1

- The student should mix the copper oxide with the sulfuric acid. Keep on adding copper oxide until it no longer dissolves and there is some of the powered left in the bottom of the beaker. The solution should then be left to crystallise for one day.

- Is it a good answer?

Is it a good answer?

Good

- They have described the reaction.
- They have explained how the copper sulfate crystals will be formed.

Could be improved

- The answer lacks detail.
- They have missed out some steps.

Student response 1: Commentary

- The student should **mix the copper oxide with the sulfuric acid** (1).

They have correctly described the first stage of the process.

- Keep on **adding copper oxide until it no longer dissolves and there is some of the power left** (1) in the bottom of the beaker.

They have not used the word ‘excess’ but the description is clear enough for a mark.

Student response 1: Commentary

- The solution should then be **left to crystallise** (1) for one day.

They have described the final step of the process but they have forgotten to describe the excess copper oxide being filtered out or the need to heat up the solution to remove some of the water. Overall they achieved 3 of the 5 possible marks.

Student response 2

- Add an excess copper oxide to the sulfuric acid. Heat the mixture slightly to allow more copper oxide to dissolve. Allow the solution to cool and then filter out any remaining copper oxide using filter paper and a filter funnel. Heat the solution again so that some of the water evaporates off to make a more concentrated copper sulfate solution. Allow the solution to crystallise in a warm oven.

This is a strong answer.

Student response 2: Commentary

- Add an **excess copper oxide to the sulfuric acid** (2). Heat the mixture slightly to allow more copper oxide to dissolve.

The process is clearly explained gaining the student 2 marks: one for describing missing the reactants; one for stating that copper oxide should be in excess.

- Allow the solution to cool and then **filter out any remaining copper oxide using filter paper** (1) and a filter funnel.

The removal of the excess copper oxide after the reaction has taken place is described well here gaining another mark.

Student response 2: Commentary

- **Heat the solution again so that some of the water evaporates off** (1) to make a more concentrated copper sulfate solution.

A good description of the next step and the reason for it.

- Allow the solution **to crystallise in a warm oven** (1).

The final step of the process, obtaining the copper sulfate crystals, is described. Stating that the solution should be left to crystallise would also have been good enough to gain the mark. The student gained full marks.

Mark scheme

Answer	Marks
1. Mix copper oxide and sulfuric acid (heat)	1
2. Added copper oxide in excess	1
3. Filter out excess copper oxide	1
4. Heat solution to remove some of the water	1
5. Leave to crystallise	1