Metals

Teacher Notes

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Metals

Why focus on G&T and higher achievers?

Within the education system every child has the right to develop their learning so as to maximise their potential.

These exercises are designed to give students enthuse and enrich activities that although related to the curriculum are in fact taking the learning experience to the next level whilst also showing chemistry in a familiar context. This has been found to be a successful model for not only improving learning but also for raising levels of motivation. Higher achieving students can find the restraints of the standard curriculum to be demotivating leading to underachievement.

The different activities are designed to improve a number of skills including practical work/dexterity, thinking/analysis skills, literacy, research activities, use of models and teamwork. Students should also gain confidence through the activities and improve the ability to express themselves.

Some of the activities would appear to be complex for KS3 (year 9), however at this stage in their learning high achieving students are open to new concepts and are ready to explore issues without pre-conceptions. They are keen to link ideas and develop concepts and understanding. It can prove to be an uplifting experience.

Introduction

Understanding the properties and reactivity of metals can help to develop students basic understanding of chemical reactions. All the properties of a particular substance depend upon the elements present and how they are bonded to each other. The activities develop such areas as elements, ions, reactivity series, metal extraction, displacement reactions and oxidation/reduction.

This programme is also designed to develop students thinking, investigative and research skills.

Topic	Type of activity	Summary	Timing (mins)	KS3	KS4	KS5	Page
Alkali metals	Demonstration	A demonstration to illustrate the extreme reactivity of metals, their varying properties and the beginning of a reactivity series.	20	√	√		6
Iron as a metal	Thinking skills, misconceptions	A series of true or false questions focused around metallic iron	Student centred	√	√		10
The reduction of iron oxide by carbon	Practical	This activity reinforces the idea of reactivity, in that carbon displaces iron from iron(III) oxide. It also links to a context- the extraction of iron from its ore. Ideas about redox can also be explored.	30	√	√	√	15
Displacement reactions:	Practical	Further building on the ideas of reactivity, this activity focuses on	30	√	√	√	19

metals & their salts		the displacement of metals from a solution of metal salts. It extends the series from activity 1.					
Thermite reaction	Demonstration	This again shows reactivity within a context-the displacement of iron from iron(III)oxide by aluminium. A process used to weld railway tracks.	30	√	√	√	21
Reactivity	Thinking skills, misconceptions	A series of multiple choice questions about the electronic structure of atoms and ions.	Student centred		√	√	24
Rusting	Lateral thinking exercise	This activity encourages students to think laterally in order to solve a series of problems related to rusting on ships. It develops ideas regarding corrosion and its prevention.	Student centred	√	√		31
Diagnostic test	Summative assessment	An activity exploring the level of knowledge and understanding. Revision for key stage 5.	50-75	√	√	√	31

The first activity is an introduction to the varying properties and reactivity of metals, focussing on the more reactive, and therefore dramatic, examples. Although it is easy to see a difference between the reactions of the metals with water, it is important to also compare observations of hardness, malleability and even the differing colour of the oxide layer. Students should focus on describing the reaction before explaining the chemistry, two skills often confused on examination papers.

This feeds into the second activity where another metal, iron, is taken as an example in order to probe students understanding of metals and identify, and rectify, any misconceptions that are thrown up. Although designed for key stage 4 it is suitable for Year 9 students.

This leads into an activity that reinforces the idea of reactivity, in that carbon displaces iron from iron(III)oxide. It also links to a context-the extraction of iron from its ore. Ideas about redox can also be explored as well as providing an opportunity to develop equation writing skills.

The varying reactivity of metals is then exploited by investigating the displacement of metals from solutions of their salts. This enables students to extend the reactivity series that they started in Activity 1 to include extra, perhaps more familiar, metals. The thermite reaction allows for more contextualised learning with an example of the varying reactivity of two metals being used to produce molten iron in order to weld together railway tracks. It can also be used to further develop the idea of redox.

Exploring the chemistry behind reactivity can be a powerful learning tool as it can then be related to unrelated reactions. Although Sctivity 6 is designed for key stage 5 students it is reasonable to use it with year 11 students who have by that stage a broad understanding of chemistry.

As a means of developing higher order thinking skills alongside lateral and creative thinking the rusting exercise is a good resource. It introduces concept maps as a way of presenting ideas. This is a teacher led exercise and includes powerpoint presentations, worksheets, handouts and experiments. It can also be used to promote discussion.

Finally the diagnostic test can be used as a summative assessment tool to provide an insight into the level of knowledge and understanding of this topic.

Aims and objectives

The aims and objectives of these activities are:

- Developing questioning skills through problem solving.
- Exploring the use of models to expand understanding
- To appreciate that all models have limitations dependent upon the knowledge available and the application.
- Develop practical skills and dexterity.
- Promote independent learning and research skills.
- Chemistry topics:
    - Word/Symbol equations
    - Properties of metals
    - Reactivity series
    - Displacement reactions
    - Redox
    - Corrosion

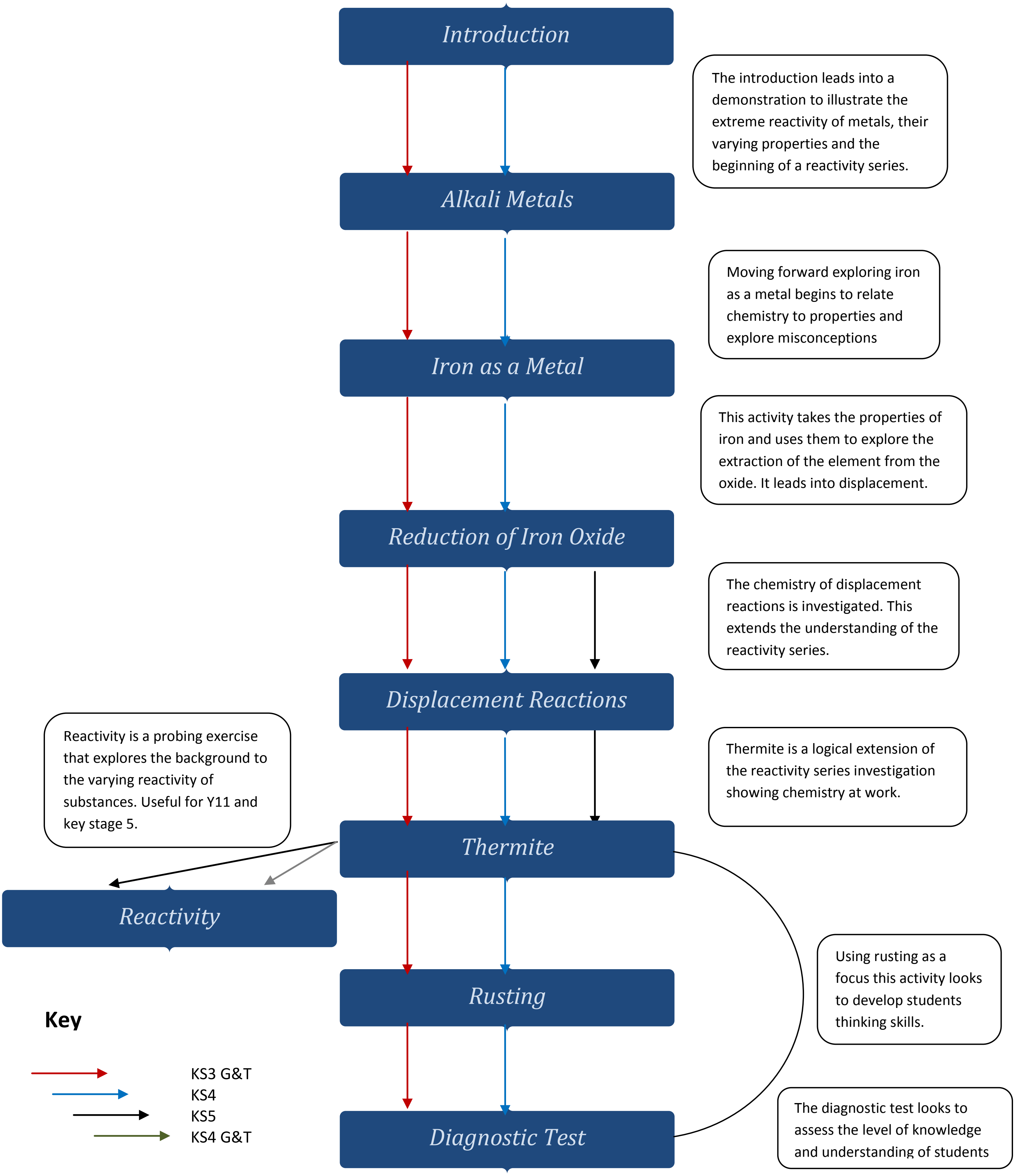
These exercises can be used with key stages 3 and 4, as indicated on the *Possible Routes*. They can also be use as revision exercises for key stage 5.

These activities have proved very successful with key stages 3 and 4 and have stimulated students into further independent learning. They have enhanced understanding of metals, their properties, extraction and uses.

These exercises provide a reinforcement and revision tool for a number of topics from the A level syllabus.

At all levels there is promotion of questioning skills, independent learning and research skills.

Possible routes

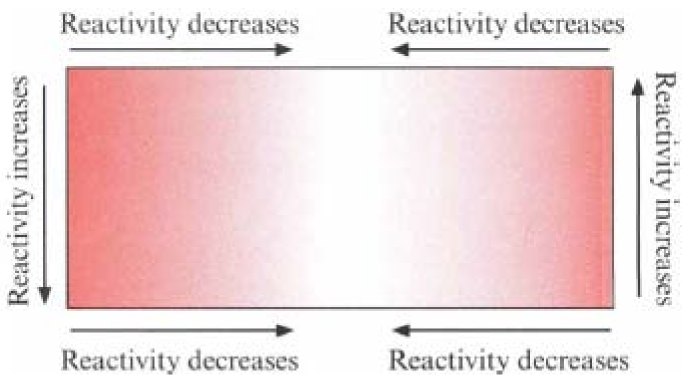


Trends in the Periodic Table

Elements in the groups on the left-hand side of the periodic table tend to react by losing electrons to form positive ions. As you go down a group, it gets easier for atoms to lose electrons. In other words, as you go down a group of metals it gets easier for the elements to react; they become more reactive. Another way of saying this is that reactivity increases as you go down a group on the left-hand side of the table.

Similarly, as you go across a period from left to right, atoms have to lose more electrons to form stable positive ions. The more electrons have to be lost, the more energy it takes, so the harder it gets to form positive ions. In other words, as you go from left to right across a period on the lefthand side of the table, reactivity decreases.

In the same way, on the right-hand side of the table elements tend to react by gaining electrons. As you go down a group, it gets harder to gain electrons and reactivity decreases. Similarly, as you go from right to left across a period, more electrons have to be gained which makes it harder, so reactivity decreases. So, on the right-hand side of the table, reactivity increases up the group, and decreases from right to left in a period.



General reactivity trends in the periodic table

Reactivity Series

If you look at Group 2 in the periodic table, you will see that calcium falls below magnesium; calcium is therefore more reactive than magnesium. Sodium falls to the left of magnesium, in the same period; sodium is that calcium and sodium will be fairly alike in terms of reactivity. This is, indeed, the case. They both, for example, react quite vigorously with cold water, corrode rapidly in air and burn when heated in air. Sodium, however, reacts more vigorously than calcium. We can therefore arrange these three elements in order of reactivity thus:

Sodium (most reactive)
Calcium
Magnesium (least reactive)

A good demonstration to illustrate this is given below. The demonstration is useful for teaching students to be clear on describing what they see and then using their observations to develop a hypothesis for the chemistry involved.

Activity 1: Alkali metals

These demonstrations show the similarity of the physical and chemical properties of the alkali metals and the trend in reactivity down Group 1 of the Periodic Table.

Lesson organisation

These experiments must be done as a demonstration at all levels of school chemistry.
The experiments take about 10 – 20 mins if everything is prepared in advance.
Advance preparation includes cutting pieces of alkali metals to the recommended size, filling water troughs and setting up safety screens.

Technical notes

Lithium (**Highly flammable, Corrosive**) Refer to CLEAPSS Hazcard 58
Sodium (**Highly flammable, Corrosive**) Refer to CLEAPSS Hazcard 88
Potassium (**Highly flammable, Corrosive**) Refer to CLEAPSS Hazcard 76
Ethanol (**Highly flammable**) Refer to CLEAPSS Hazcard 40A
or Industrial denatured alcohol (**Highly flammable, Harmful**) Refer to CLEAPSS Hazcard 40A
2-methylpropan-2-ol (**Highly flammable, Harmful**) Refer to CLEAPSS Hazcard 84B
Universal indicator solution (**Highly flammable**) Refer to CLEAPSS Hazcard 32 and Recipe book sheet 47

1 A technician should prepare the pieces of metal and store them under oil. Using the tweezers, remove a large piece of the alkali metal from the oil. Ensure that conditions are dry. Place the metal on a tile and, using a scalpel or sharp knife, cut pieces of lithium (5 mm cubes), sodium (4 mm cubes) and potassium (3 mm cubes). Place the small pieces in separate bottles of oil, labelled with the metal name and the hazard symbol. Cut each alkali metal separately and return the larger piece to its bottle before starting the next one.

2 Place any apparatus used to cut (and later handle) the metal (filter paper, scalpels etc) in a trough of water after use. Small pieces of alkali metal for disposal should be allowed to react fully with ethanol (for lithium and sodium) or 2-methylpropan-2-ol (for potassium) until fizzing stops, before washing away with water.

Teaching notes

The reaction with water can be done with the trough on an overhead projector. The projector should first be focussed on a matchstick floating in the water. Alternatively use a flexicam linked to a projector or a TV screen.

The reactions are:

$$2M(s) + 2H_2O(l) \rightarrow 2MOH(aq) + H_2(g)$$

where M represents the alkali metal. The solutions of the hydroxides formed are alkaline.

The reactions clearly show that the reactivity sequence is lithium< sodium < potassium

Thus, by looking at the positions of the elements in the periodic table and their reactions we can build up a reactivity series, with the most reactive elements at the top and the least reactive at the bottom. We can include as many or as few elements as we wish in this series. If we only include the most common elements, we arrive at a list which looks like this:

Potassium (K)
Sodium (Na)
Calcium (Ca)
Magnesium (Mg)
Aluminium (Al)
Carbon (C)
Zinc (Zn)
Iron (Fe)
Lead (Pb)
Hydrogen (H)
Copper (Cu)
Silver (Ag)
Gold (Au)

Such a series is very useful in explaining and predicting the outcomes of various reactions.

One example is the displacement reaction. Aluminium is more reactive than iron; to put it another way, aluminium forms positive ions more easily than iron does. Iron(III) oxide is an ionic compound: it contains positive iron ions and negative oxygen ions, If you mix aluminium metal with iron oxide, and provide some energy by heating the mixture, the aluminium atoms will give up electrons and force them on to the iron ions, thus converting the iron oxide into metallic iron:

$2Al + 2Fe^{3+} \rightarrow 2Al^{3+} + 2Fe$

To put it another way:

$2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

In other words, a more reactive element will displace a less reactive one from its compounds. If, for example, you dip a copper wire into silver nitrate solution and leave it for a while, crystals of silver will start to grow on the copper wire. The copper is more reactive than the silver and forces electrons on to the silver ions, thus converting them into metallic silver atoms:

$Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$

or

$Cu + 2Ag^{+} \rightarrow Cu^{2+} + 2Ag$

In the same way, any metal below carbon in the reactivity series can be extracted from its ore by heating the ore with carbon, because the carbon will displace the metal. One of the most important of these reactions, is the production of iron by heating iron ore (iron oxide) with coke (which is a form of carbon). The carbon, being more reactive than iron, displaces it thus:

$3C + Fe_2O_3 \rightarrow 3CO + 2Fe$

Again, you can tell that copper, silver and gold will not be attacked by *dilute* acids. This is because they are all less reactive than hydrogen and will thus not displace hydrogen from an acid. On the other hand, any metal above hydrogen in the reactivity series will react with a dilute acid.

If, for example, you add magnesium to dilute hydrochloric acid, the magnesium will displace the hydrogen from the acid as follows:

$Mg + 2HCl \rightarrow MgCl_2 + H_2$
Or
$Mg + 2H^+ \rightarrow Mg^{2+} + H_2$

You can see how important the reactivity series is in predicting the outcome of displacement reactions.

There are other uses as well; *e.g.* some compounds, when heated, will decompose. This is called thermal decomposition. Mercury oxide is a good example; it is made up of positive mercury ions and negative oxygen ions (Hg^{2+} and O^{2-}). Mercury is very unreactive: it comes between copper and silver in the reactivity series. It was reluctant to give up its electrons in the first place to form positive ions. If you heat mercury oxide, the oxygen ions give back electrons to the mercury ions. The mercury ends up as metallic mercury, and the oxygen ions end up as neutral, gaseous oxygen molecules:

$2HgO \rightarrow 2Hg + O_2$
Or
$2Hg^{2+} + 2O^{2-} \rightarrow 2Hg + O_2$

On the other hand, if you heat sodium oxide, nothing happens. This is because sodium is towards the very top of the reactivity series. It was eager to form positive ions in the first place by giving away electrons, so sodium ions are less likely than mercury ions to accept the electrons back again and form neutral atoms. The general rule is that the more reactive a metal is, the more stable its compounds are likely to be, *i.e.* the less likely to be split up by heating.

Because the reactivity series is so important as a tool for thinking about the outcome of reactions, it is worthwhile trying to remember it. The following is a mnemonic for the series, using only the elements in the series printed out above, *i.e.* if you remember this sentence, each word starts with the same letter as the corresponding element. You can either use this mnemonic or make up your own: you can use an extended reactivity series, putting in other elements. The general rule for mnemonics is that the more ridiculous the picture, the better they are likely to be at reminding you of the order.

Perhaps	Potassium
Some	Sodium
Cows	Calcium
May	Magnesium
All	Aluminium

Come	Carbon
Zooming	Zinc
In	Iron
Large	Lead
Herds	Hydrogen
Chewing	Copper
Some	Silver
Grass	Gold

You will notice that almost all the elements in this series are metals. The same sorts of arguments work for the non-metals. The halogens provide a good example. The order of reactivity for these is the same as the order of the elements in the group:

Fluorine
Chlorine
Bromine
Iodine
Astatine

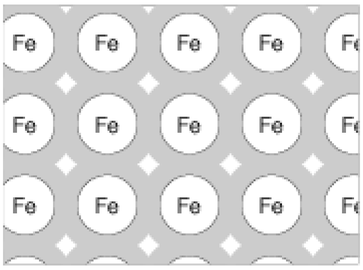
Thus, for example, if you pass chlorine gas through potassium bromide solution, the solution will go reddish-brown because the chlorine displaces the bromine from the potassium bromide:

$2KBr + Cl_2 \rightarrow 2KCl + Br_2$
or
$2Br^- + Cl_2 \rightarrow 2Cl^- + Br_2$
The higher element in the reactivity series displaces the lower one from a solution of its salt.

Activity 2: Iron as a Metal

Iron – true or false?

The statements below refer to the diagram of the structure of iron. The diagram shows part of a slice through the three dimensional structure.



Please read each statement carefully, and decide whether it is correct or not.

1. Iron has a type of bonding called metallic bonding. **TRUE/FALSE**

2. Iron atoms do not have a full outer shell of electrons, and this makes iron very reactive.
TRUE/FALSE

3. An iron atom is a silver-grey colour, and so iron metal is a silver-grey colour.
TRUE/FALSE

4. Iron can conduct electricity because some of the iron atoms can slip over their neighbours, and move through the solid.
TRUE/FALSE

5. Iron can be reshaped, without changing the shape of iron atoms.
TRUE/FALSE

6. The reason iron rusts is that iron atoms will rust if exposed to damp air.
TRUE/FALSE

7. In iron metal each atom is bonded to each of the other iron atoms surrounding it.
TRUE/FALSE

8. Iron conducts electricity because iron atoms are electrical conductors.
TRUE/FALSE

9. Iron is a solid because that is the natural state for metals.
TRUE/FALSE

10. A metal such as iron consists of positive metal ions, and negative electrons which move around the solid between the ions.
TRUE/FALSE

11. An iron atom will reflect light, and so freshly polished iron shines.
TRUE/FALSE

12. The reason that iron becomes a liquid when heated is because the bonds melt.
TRUE/FALSE

13. Iron conducts electricity because it contains a ‘sea’ of electrons.
TRUE/FALSE

14. The atoms in a metal such as iron are held together by ionic bonds.
TRUE/FALSE

15. The reason iron conducts heat is because there is room between the atoms for hot air to move through the metal.
TRUE/FALSE

16. The reason that iron is hard is because iron atoms are hard.
TRUE/FALSE

17. In iron there are molecules held together by magnetism.
TRUE/FALSE

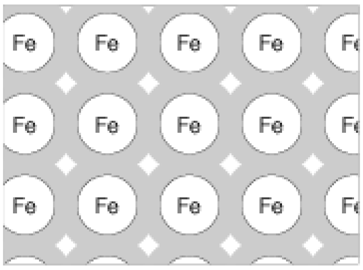
18. If a metal such as iron is heated to a very high temperature it would become a gas.
TRUE/FALSE

19. Metals such as iron expand when heated because the atoms get bigger.
TRUE/FALSE

20. Chemical bonds are needed to hold the atoms together in a metal such as iron, even though all of the atoms are of the same type.
TRUE/FALSE

Iron – answers

Below you will find listed the 20 statements you were asked to think about. Following each is a brief comment suggesting whether or not the statement is true, and why.



1. Iron has a type of bonding called metallic bonding. **TRUE**: iron is a metal, and all metals have a type of bonding called metallic bonding which is different from covalent and ionic bonding. In metallic bonding the outer shells of adjacent atoms overlap, and the outer shell electrons are free to move about through the lattice. The metal consists of metal cations and a balancing number of these 'free' electrons.

2. Iron atoms do not have a full outer shell of electrons, and this makes iron very reactive. **FALSE**: although an isolated iron atom has an electronic configuration of 2.8.14.2, the outer electrons are involved in the bonding in the metal. Iron is not very reactive, although it will slowly rust.

3. An iron atom is a silver-grey colour, and so iron metal is a silver-grey colour. **FALSE**: the colour of iron is a property of the arrangement of cations and electrons. A single atom of iron would not have a colour.

4. Iron can conduct electricity because some of the iron atoms can slip over their neighbours, and move through the solid. **FALSE**: the iron cations are normally fixed in their lattice positions. It is the electrons from the outer shells that are able to move about, allowing electrical current to flow through the metal.

5. Iron can be reshaped, without changing the shape of iron atoms. **TRUE**: metals can be worked into different shapes by hammering to force the cations to slip over each other. The cations change position, but not shape.

6. The reason iron rusts is that iron atoms will rust if exposed to damp air. **FALSE**: the rusting of iron is due to a chemical reaction between the iron and oxygen and water vapour in the air. During these reactions some of the iron cations and electrons become part of a new chemical compound (the rust), but the atoms themselves do not corrode.

7. In iron metal each atom is bonded to each of the other iron atoms surrounding it. **TRUE**: the iron atoms are packed together so that each iron cation is surrounded by eight others as if it is in the centre of a cube. The structure is held together by metallic bonding.

8. Iron conducts electricity because iron atoms are electrical conductors. **FALSE**: the metal conducts because some electrons are able to move through the metallic lattice structure. The individual atoms can not be considered to conduct. The outer electrons are only able to leave the cations because the outer electron shells overlap.

9. Iron is a solid because that is the natural state for metals. **FALSE**: the natural state depends on the temperature. Deep in the earth – where it is very hot – iron is a liquid. One metal called mercury is a liquid at room temperature.

10. A metal such as iron consists of positive metal ions, and negative electrons which move around the solid between the ions. **TRUE**: the iron structure contains iron cations surrounded by the fast moving electrons that would be in outer shells of separate iron atoms. (Sometimes this is called a 'sea of electrons'.)

11. An iron atom will reflect light, and so freshly polished iron shines. **FALSE**: polished metal will form a mirror because of the regular lattice of cations and the 'sea' of electrons. Individual iron atoms would not reflect light.

12. The reason that iron becomes a liquid when heated is because the bonds melt. **FALSE**: the metal melts when enough energy is provided to allow the cations to slip over each other. The bonds in the liquid metal are weaker than in a solid metal. If the liquid was heated until it boiled the bonds would break (but not 'melt').

13. Iron conducts electricity because it contains a 'sea' of electrons. **TRUE**: the electrons are able to move about, and will pass along the metal when it is connected to a battery.

14. The atoms in a metal such as iron are held together by ionic bonds. **FALSE**: the bonding in a metal is metallic bonding. This is different from ionic bonding as there are no anions (negative ions) present.

15. The reason iron conducts heat is because there is room between the atoms for hot air to move through the metal. **FALSE**: the iron cations are held close together by the metallic bonding, and there is no room for other atoms and molecules to get between them. Heat passes along the metal due to lattice vibrations and the movement of electrons.

16. The reason that iron is hard, is because iron atoms are hard. **FALSE**: hardness is a property of the metal due to the strong bonding holding the structure together. It is the arrangement of cations and free electrons which makes the metal hard.

17. In iron, there are molecules held together by magnetism. **FALSE**: there are no molecules in a metal - each cation is bonded to all those around it by the 'sea' of electrons, and those cations are bonded to others, and so on. Each cation in a metallic crystal is bonded (indirectly) to all the others.

18. If a metal such as iron is heated to a very high temperature it would become a gas. **TRUE**: if a solid metal is heated it will melt, and if heating is continued to a high enough temperature the liquid metal will boil.

19. Metals such as iron expand when heated because the atoms get bigger. **FALSE**: when the metal is heated the cations vibrate more, and move a little further apart.

20. Chemical bonds are needed to hold the atoms together in a metal such as iron, even though all of the atoms are of the same type. **TRUE**: the atoms would not remain joined together if there was no bonding between them. This is true for all solids whether the atoms are of one type (in an element) or several (in a compound).

Displacement Reactions

Displacement reactions are characterised by one element displacing another out of its compound.
Iron used to be extracted from its ore, haematite, which is chemically iron(III) oxide, by heating the ore with charcoal. Nobody really knows how the first metals were extracted from their oxides, thousands of years ago, but it could have been by an accidental reduction of the oxide rock being heated in a charcoal fire and then the liquid metal solidifying. The wood and charcoal which were burning contain carbon, and carbon is more reactive towards oxygen than iron. Because it had a greater affinity for the oxygen, the carbon (given the energy from the fire) can take up the oxygen and leave iron on its own. Another way of saying this is that carbon can displace iron from its oxide.

- The carbon combines with the oxygen out of the iron oxide and keeps it for itself
    - The carbon is oxidised *(i.e.* picks up oxygen)
- The iron oxide loses oxygen to the carbon
    - The iron oxide is reduced. (Reduction is the loss of oxygen.)

Whenever one element is reduced (loses oxygen), another must be oxidised (gain oxygen). This, as with other types of chemical reactions, can be classified in different ways depending on what you are looking at; here the reaction can be considered as a displacement reaction or a redox reaction. This reaction can be scrutinised for oxidation number changes. Iron goes from +3 to 0 and so has been reduced, and the carbon is oxidised from zero oxidation number to **+** 2 in carbon monoxide.

$Fe_2O_3(s) + 3C(s) + 2Fe(s) + 3CO(g)$

(Actually, in the blast furnace it is carbon monoxide which does the greatest amount of the reduction.)

A web link to a useful resource looking at industrial blast furnaces (video, resources and data) is give below:

http://www.rsc.org/Education/Teachers/Resources/Alchemy/index2.htm

There is also a very quick class practical illustrating the reaction whereby iron is produced from iron(III)oxide on a match head.

Activity 3: The reduction of iron oxide by carbon

Teacher notes

Topic

Reactivity series, oxidation and reduction.

Timing

20–30 min.

Description

Students react iron(III) oxide with carbon on the end of a spent match to produce iron.

The mixture is fused with sodium carbonate.

Apparatus and equipment (per group)

- Match (non-safety) (see note 1)
- Spatula
- Bunsen burner
- Magnet.

Chemicals (per group)

- Sodium carbonate (access to a small quantity to adhere to the end of a match)
- Iron(III) oxide (access to a small quantity to adhere to the end of a match).

Teaching tips

A fragment of metallic iron will be attracted to the magnet. Sodium carbonate fuses easily and brings the oxide into intimate contact with the carbon from the burnt match head. One suggestion is that students test iron(III) oxide with a magnet before the reaction to show it is not magnetic.

Background theory

Reactivity series, reduction.

Health & Safety

Wear eye protection.

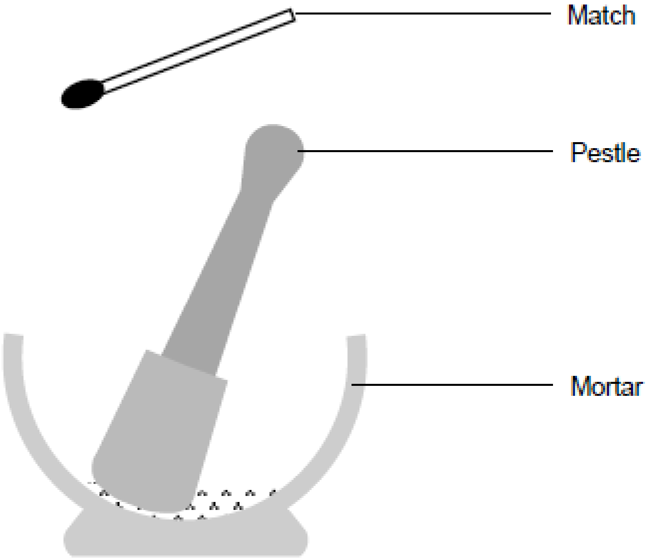
Answers

1. Reduction can be described as the removal of oxygen.
2. Carbon comes between aluminium and iron in the reactivity series. The products of the reaction between carbon and zinc oxide.
3. Calcium is reactive and therefore its oxide is very stable, carbon is not reactive enough to displace oxygen from calcium oxide.

The reduction of iron oxide by carbon

Introduction

Metals high in the reactivity series will reduce the oxides of those lower in the series. The oxides of metals between zinc and copper in the reactivity series can be reduced by carbon. In this experiment, sodium carbonate is used to fuse the reactants in intimate contact.



What to do

1. Char the point of a used match, moisten it with a drop of water and rub on some sodium carbonate crystals.

2. Rub the point in some powdered iron(III) oxide (Fe_2O_3) and heat in a blue Bunsen burner flame until the point glows strongly.

3. Allow to cool.

4. Crush the charred head in a mortar and pestle then run a magnet through the pieces.

Health & Safety

Wear eye protection.

Questions

1. What does 'reduction' mean?

2. Carbon does not reduce aluminium oxide. Where would carbon be placed in this reactivity series?

Potassium
Sodium
Calcium
Magnesium
Aluminium
Zinc
Iron
Lead
Copper
What other information would you need to determine carbon's exact place?

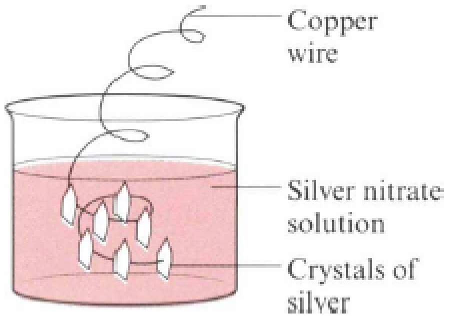
3. Explain why calcium oxide cannot be reduced using carbon.

You can arrange the metals in order of reactivity. The elements at the top are the most reactive. Anything nearer the top will displace anything lower down from its compounds. (Note that you should never attempt a displacement reaction with any of the top five elements – it is far too dangerous!) This list is called a reactivity series.

To show that the elements nearer the top will displace the elements lower down from its compounds, when you dip some copper wire into a solution of silver nitrate, crystals of silver will start to grow on the copper wire because the copper, being more reactive than silver, will displace silver from the silver nitrate. The solution will go blue, the usual colour for copper compounds in solution.

This can make a good demonstration in winter if a sheet of copper is cut into the shape of a conifer. The silver formed looks like ice and snow clinging to the leaves and branches!

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$



Copper displaces silver from silver nitrate solution

The reverse will not work. Silver will not displace copper from copper sulfate solution, and nothing lower down the series will displace an element higher up the series. So copper will never displace hydrogen gas from a solution of a dilute acid, but zinc will.

The majority of the elements in the list are metals, but carbon and hydrogen are included because of the importance of carbon in extracting metals, and because of the importance of metal/acid reactions.

Any metal which is below carbon (coke) in the series can be displaced by carbon from its oxide.

Any metal above carbon (such as aluminium) cannot be extracted with coke.

Sodium can only be extracted by electrolysing molten salt (sodium chloride). The top four metals in the reactivity series shown here were not discovered until electrolysis of their molten compounds became possible early in the nineteenth century.

Non-metals also vary in reactivity. In, for example, Group 17 of the periodic table (the halogens), chlorine is more reactive than iodine. If you pass chlorine gas into potassium iodide solution, the solution goes brown because the chlorine displaces iodine from the potassium iodide. Iodine in KI has an oxidation number of -1 and this rises to 0 in the elemental iodine, so iodine has been oxidised. The opposite occurs with chlorine.

$2KI(aq) + Cl_2(aq) \rightarrow 2KCl(aq) + I_2(aq)$

Practicals that can be used to show the principals and uses of displacement include a class practical looking at the displacement of metals from solutions of their ions and the much requested thermite reaction. The latter is still used to weld together railway lines.

Activity 4: Displacement reactions between metals and their salts

Some metals are more reactive than others. In this experiment, a strip of metal is added to a solution of a compound of another metal. A more reactive metal displaces (pushes out) a less reactive metal from its compound. In carrying out the experiment, students investigate competition reactions of metals and arrive at a reactivity series of the four metals they use.

Lesson organisation

There are many ways of carrying out this series of reactions. The one described here uses a spotting tile but the same procedure could be adapted for use with test -tubes. The advantages of the spotting tile method include:

- Very small quantities of chemicals are used.
- The whole set of experiments is displayed together, making comparison easier.
- Clearing-up afterwards is simple and avoids metal deposits being left in sinks.

Careful thought needs to be given to distribution of the chemicals to the class. Solutions could be distributed in test-tubes, or in small bottles fitted with droppers for sharing between several pairs of students. Metals could be issued in sets. The teacher should keep control of the magnesium ribbon, dispensing short lengths when required.

There should be no flames alight so that students are not tempted to burn pieces of magnesium and the teacher should be alert to the possibility of pieces of magnesium being removed from the laboratory.

The experiment should take about 30 minutes.

Apparatus and chemicals

Eye protection

Each student or pair of students will require:

Spotting tile, with at least 16 depressions (or two smaller tiles)
Dropping (teat) pipette
Beaker (100 cm^3)
Felt tip pen or other means of labelling

Access to about 5 cm^3 each of the following 0.1 mol dm^{-3} metal salt solutions:

Zinc sulfate (**Low Hazard** at this concentration). Refer to CLEAPSS Hazcard 108.
Magnesium sulfate (**Low hazard**). Refer to CLEAPSS Hazcard 59B.
Copper(II) sulfate (**Low Hazard** at this concentration). Refer to CLEAPSS Hazcard 27B.
Lead(II) nitrate (**Toxic**, **Dangerous for the environment**). Refer to CLEAPSS Hazcard 57A.

Five samples, approximately 1 cm lengths or squares, of the following metals. The metals, except lead, present are low hazard as used here.
Zinc foil. Refer to CLEAPSS Hazcard 107.
Magnesium ribbon. Refer to CLEAPSS Hazcard 59A.
Copper foil. Refer to CLEAPSS Hazcard 26.
Lead foil (**Toxic**, **Dangerous for environment**). Refer to CLEAPSS Hazcard 56.

Technical notes

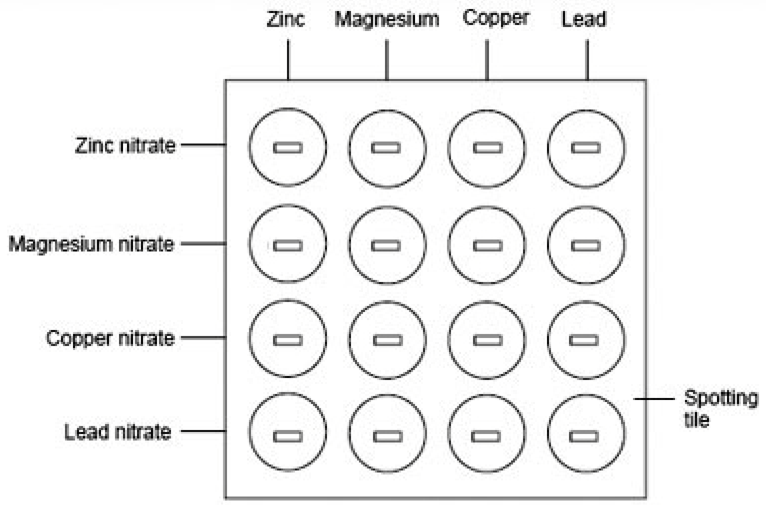
Zinc sulfate (**Harmful**, **Oxidising**) Refer to CLEAPSS Hazcard 108.
Magnesium sulfate (**Low Hazard**) Refer to CLEAPSS Hazcard 59B.
Copper(II) sulfate (**Harmful**) Refer to CLEAPSS Hazcard 27B.
Lead nitrate (**Toxic**, **Dangerous for the environment**) Refer to CLEAPSS Hazcard 57A.
Zinc foil. Refer to CLEAPSS Hazcard 107.
Magnesium ribbon. Refer to CLEAPSS Hazcard 59A.
Copper foil. Refer to CLEAPSS Hazcard 26.
Lead foil (**Toxic**, **Dangerous for Environment**) Refer to CLEAPSS Hazcard 56.

1 Solutions may be dispensed in 5 cm^3 beakers to each pair of students or in small bottles fitted with droppers to groups of students.

2 Metals should be approximately 1 cm lengths or squares of ribbon or foil cleaned with emery cloth and as similar in size as possible.

Procedure

a Using a dropping pipette, put a little of the zinc nitrate solution in four of the depressions in the spotting tile, using the following illustration as a guide. Label this row with the name of the solution. Rinse the pipette well with water afterwards.



b Do this for each solution in turn , rinsing the pipette when you change solution.

c Put a piece of each metal in each of the solutions, using the illustration as a guide.

d Over the next few minutes observe which mixtures have reacted and which have not.

Teaching notes

Remind the class that they are looking for cases where one metal displaces another. Some of the solutions are slightly acidic so that bubbles of hydrogen are sometimes seen. Explain that this does not count as displacement of one metal by another.

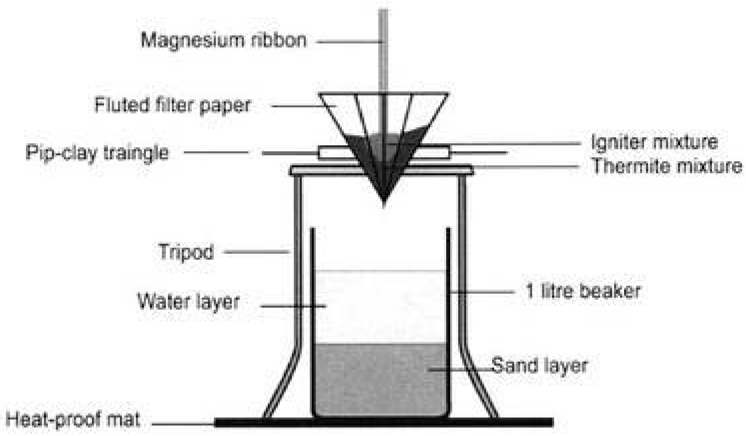
It might be best to get the class to tell you what they think the order of reactivity is while they still have the evidence in front of them, so that apparent discrepancies can be resolved.

Activity 5: The thermite reaction

This demonstration shows the highly exothermic reaction between aluminium and iron(III) oxide that produces molten iron. This is a competition reaction, showing aluminium to be a more reactive metal than iron. A redox reaction takes place.

Procedure

HEALTH & SAFETY: A face shield or goggles and a laboratory coat (it can become messy at the end) should be worn by the demonstrator. Do not exceed the stated quantities. See CLEAPSS guide L195 *Safer Chemicals, Safer Reactions,* section 9.4 and/or *Hazcards* 1, 11, 59A. Aluminium powder and magnesium powder/ribbon are highly flammable, barium nitrate/peroxide are oxidising and harmful.
Students should stand further than 4 m from the reaction and wear eye protection. Explosions have been known, so safety screens surrounding the apparatus are essential.



a) Fold two 12 cm diameter circles of filter paper into fluted cones and place one inside the other.
b) Into a 1 dm^3, thick-walled beaker, pour dry sand until it is one-third full and then add water until it is two-thirds full.

c) Cover an area of the bench with several heat resistant mats and place the beaker in the centre. Set up the equipment as shown in the diagram above and surround it with safety screens. Add the Thermite mixture (see note 2) to the fluted filter paper cone sitting in the pipe clay triangle.
d) Make a depression in the Thermite mixture with a spatula and place the igniter mixture (see note 4) into it.
e) Insert a magnesium ribbon fuse upright into the igniter mixture. It must extend above the fluted filter paper. Light the magnesium fuse with a Bunsen burner flame and retreat to a safe distance behind the safety screens. A very vigorous reaction should follow, with some sparks flying upwards. The very hot residue containing molten iron will fall through into the water.
f) Once the reaction has stopped, remove the beaker and decant the water into the sink. Retrieve the iron formed with a magnet. Wash the iron under running water.

Notes

The reaction is: iron(III) oxide + aluminium → aluminium oxide + iron

This shows that aluminium is above iron in the reactivity series.

The 'Thermite' mixture is stable until strong heating is applied, hence the need for an initiating reaction. Once underway, the reaction is highly exothermic, rapidly reaching temperatures as high as 2000 ^{o}C, well in excess of the melting point of iron (1535 ^{o}C). The practical use of this reaction to weld railways lines together should be mentioned – see web link below.

Web links

There are many video clips of Thermite reactions on the internet, some carried out on a scale and in a manner which is extremely hazardous.

Note that in the following reaction a much coarser mixture of the solids, as in commercial Thermite charges, is used. Using powdered solids on this scale would be extremely hazardous.
http://www.davidavery.co.uk/thermite/index.htm

Details and pictures of the thermite welding of railway tracks can be found at:
www.northeast.railfan.net/high_iron.html#thermite

Technician notes

Apparatus and chemicals

Eye protection: Safety glasses for observers, goggles or face shield for the demonstrator.

The quantities given are for one demonstration

Filter papers, 12 cm diameter, 2
Pipe-clay triangle (or similar)
Tripod
Beaker, thick-walled (1 dm^3)
Dry sand (see diagram)
Heat resistant mats
Safety screens
Small bar magnet

Thermite mixture:
Iron(III) oxide (**Low hazard**), 9 g (see notes 1 and 2)
Aluminium powder (medium grade) (**Highly flammable**), 3 g (see note 3)

Igniter mixture:
Magnesium powder (**Flammable**), 0.2 g
Barium nitrate (**Harmful**, **Oxidising**), 2 g (see note 4)

Magnesium ribbon (**Flammable**), 10 cm length

Technical notes

Aluminium powder (**Highly flammable**) Refer to CLEAPSS Hazcard 1
Magnesium powder (**Highly flammable**) Refer to CLEAPSS Hazcard 59A
Barium nitrate (**Harmful, Oxidising**) Refer to CLEAPSS Hazcard 11
Magnesium ribbon (**Low hazard**) Refer to CLEAPSS Hazcard 59A

1 It is important that the iron(III) oxide used in this demonstration is absolutely dry. An hour or so in a warm oven, or heating in an evaporating dish over a Bunsen flame, should suffice. The oxide should be allowed to cool completely before mixing.

2 The weighed quantities of iron(III) oxide (9 g) and aluminium (3 g) may be thoroughly mixed beforehand by repeatedly pouring the mixture to-and-fro between two pieces of scrap paper, and then stored for the demonstration in a container labelled 'Thermite mixture'

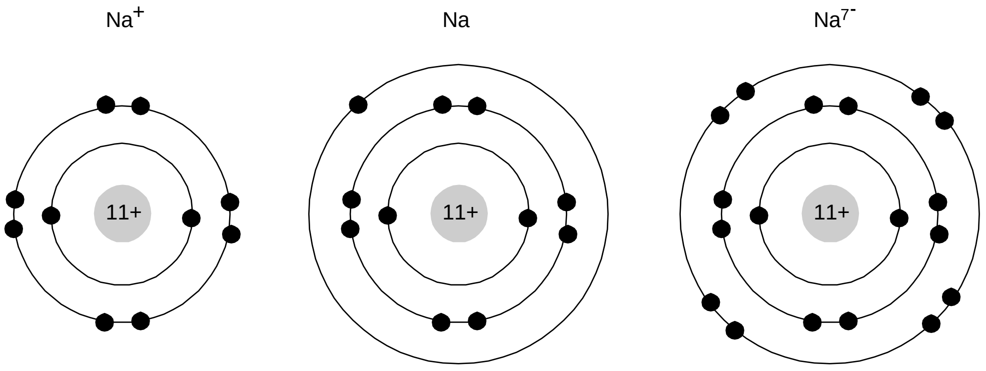
3 The demonstrator may wish (or be persuaded by the audience) to do a repeat demonstration. In this event it is important to keep the second set of materials well away from the first demonstration site.

4 The weighed quantities of magnesium powder (0.2 g) and barium nitrate (2 g) may also be thoroughly mixed beforehand as indicated in note 2, and then stored for the demonstration in a container labelled 'Igniter mixture'.

Activity 6: Reactivity

Chemical stability (1)

The diagrams below represent three chemical species:-



Sodium ion with electronic configuration of 2.8

Sodium atom with electronic configuration of 2.8.1

Sodium ion with electronic configuration of 2.8.8

1. Tick ✓ one of the four statements:

☐ Na^+ is more stable than Na

☐ Na^+ and Na are equally stable

☐ Na^+ is less stable than Na

☐ I do not know

Why did you think this was the answer?

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2. Tick ✓ one of the four statements:

☐ Na is more stable than Na^{7-}

☐ Na and Na^{7-} are equally stable

☐ Na is less stable than Na^{7-}

☐ I do not know

Why did you think this was the answer?

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3. Tick ✓ one of the four statements:

☐ Na^{7-} is more stable than Na^+

☐ Na^{7-} and Na^+ are equally stable

☐ Na^{7-} is less stable than Na^+

☐ I do not know

Why did you think this was the answer?

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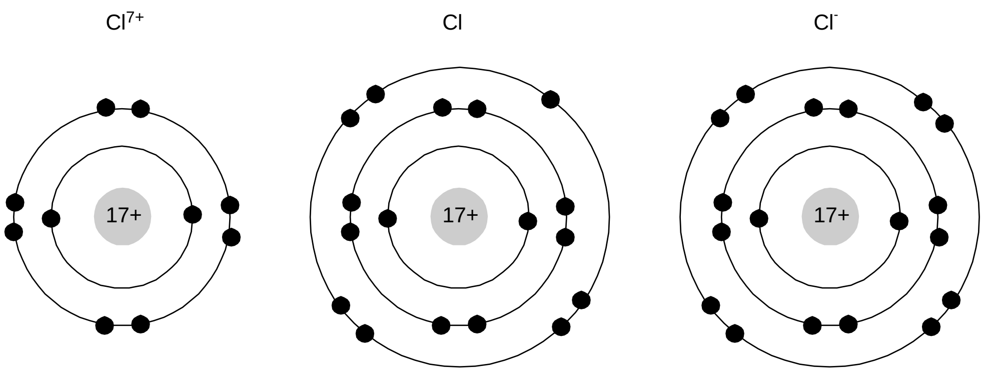
Chemical stability (2)

The diagrams below represent three chemical species:



Chloride ion with electronic configuration of 2.8

Chlorine atom with electronic configuration of 2.8.7

Chloride ion with electronic configuration of 2.8.8

4. Tick ✓ one of the four statements:

- ☐ Cl^{7+} is more stable than Cl
- ☐ Cl^{7+} is more stable than Cl
- ☐ Cl^{7+} is less stable than Cl
- ☐ I do not know

Why did you think this was the answer?

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5. Tick ✓ one of the four statements:

- ☐ Cl is more stable than Cl^{-}
- ☐ Cl and Cl^{-} are equally stable
- ☐ Cl is less stable than Cl^{-}
- ☐ I do not know

Why did you think this was the answer?

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6. Tick ✓ one of the four statements:

- ☐ Cl^{-} is more stable than Cl^{7+}
- ☐ Cl^{-} and Cl^{7+} are equally stable
- ☐ Cl^{-} is less stable than Cl^{7+}
- ☐ I do not know

Why did you think this was the answer?

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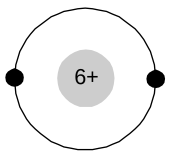
Chemical stability (3)

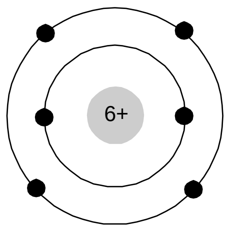
The diagrams below represent three chemical species:-

C^{4+}



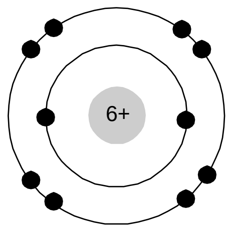
Carbon ion with electronic configuration of 2

C



Carbon atom with electronic configuration of 2.4

C^{4-}



Carbon ion with electronic configuration of 2.8

7. Tick ✓ one of the four statements:

☐ C^{4+} is more stable than C

☐ C^{4+} and C are equally stable

☐ C^{4+} is less stable than C

☐ I do not know

Why did you think this was the answer?

8. Tick ✓ one of the four statements:

☐ C is more stable than $C^{4–}$

☐ C and $C^{4–}$ are equally stable

☐ C is less stable than $C^{4–}$

☐ I do not know

Why did you think this was the answer?

9. Tick ✓ one of the four statements:

☐ $C^{4–}$ is more stable than C^{4+}

☐ $C^{4–}$ and C^{4+} are equally stable

☐ $C^{4–}$ is less stable than C^{4+}

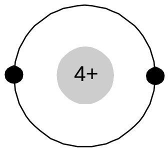
☐ I do not know

Why did you think this was the answer?

Chemical stability (4)

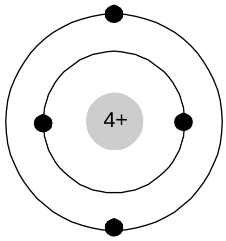
The diagrams below represent three chemical species:-

Be^{2+}



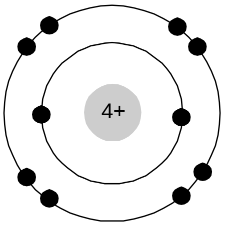
Beryllium ion with electronic configuration of 2

Be



Beryllium atom with electronic configuration of 2.2

Be^{6-}



Beryllium ion with electronic configuration of 2.8

10. Tick ✓ one of the four statements:

☐ Be^{2+} is more stable than Be

☐ Be^{2+} and Be are equally stable

☐ Be^{2+} is less stable than Be

☐ I do not know

Why did you think this was the answer?

11. Tick ✓ one of the four statements:

☐ Be is more stable than Be^{6-}

☐ Be and Be^{6-} are equally stable

☐ Be is less stable than Be^{6-}

☐ I do not know

Why did you think this was the answer?

12. Tick ✓ one of the four statements:

☐ Be^{6-} is more stable than Be^{2+}

☐ Be^{6-} and Be^{2+} are equally stable

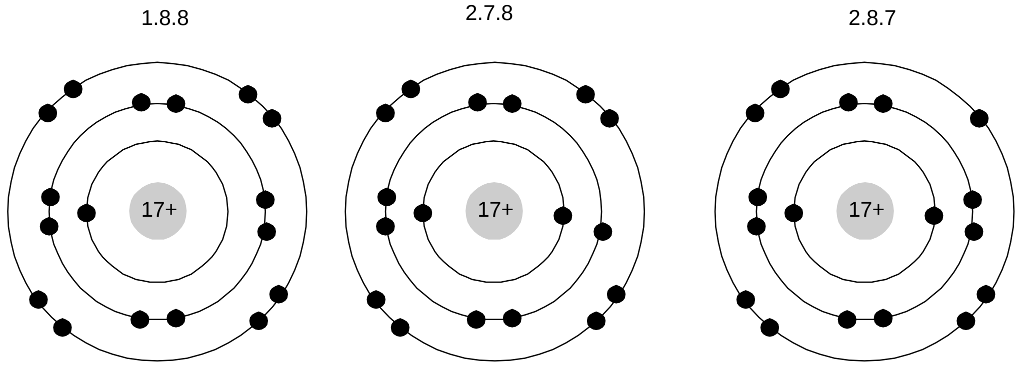
☐ Be^{6-} is less stable than Be^{2+}

☐ I do not know

Why did you think this was the answer?

Chemical stability (5)

The diagrams below represent three chemical species:-



Chlorine atom with electronic configuration of 1.8.8

Chlorine atom with electronic configuration of 2.7.8

Chlorine atom with electronic configuration of 2.8.7

13. Tick ✓ one of the four statements:

- ☐ 1.8.8 is more stable than 2.7.8
- ☐ 1.8.8 and 2.7.8 are equally stable
- ☐ 1.8.8 is less stable than 2.7.8
- ☐ I do not know

Why did you think this was the answer?

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14. Tick ✓ one of the four statements:

- ☐ 2.7.8 is more stable than 2.8.7
- ☐ 2.7.8 and 2.8.7 are equally stable
- ☐ 2.7.8 is less stable than 2.8.7
- ☐ I do not know

Why did you think this was the answer?

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15. Tick ✓ one of the four statements:

- ☐ 2.8.7 is more stable than 1.8.8
- ☐ 2.8.7 and 1.8.8 are equally stable
- ☐ 2.8.7 is less stable than 1.8.8
- ☐ I do not know

Why did you think this was the answer?

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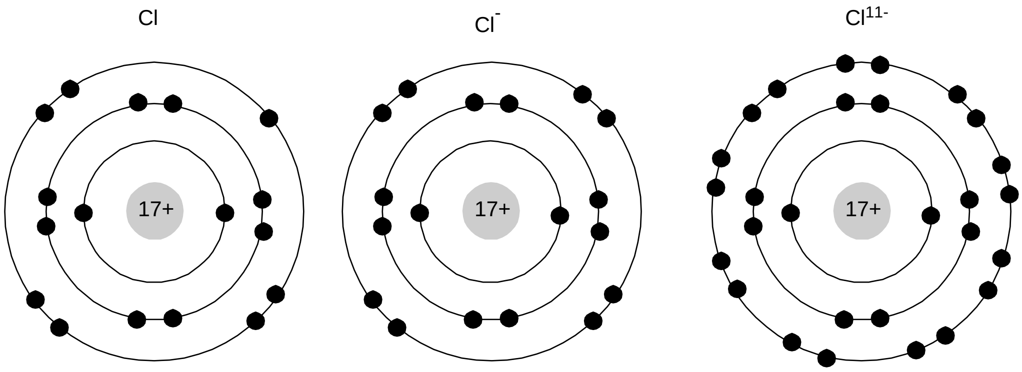
Chemical stability (6)

The diagrams below represent three chemical species:



Chlorine atom with electronic configuration of 2.8.7 | Chloride ion with electronic configuration of 2.8.8 | Chloride ion with electronic configuration of 2.8.18

16. Tick ✓ one of the four statements:

☐ Cl is more stable than Cl^-

☐ Cl and Cl^- are equally stable

☐ Cl is less stable than Cl^-

☐ I do not know

Why did you think this was the answer?

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17. Tick ✓ one of the four statements:

☐ Cl^- is more stable than Cl^{11-}

☐ Cl^- and Cl^{11-} are equally stable

☐ Cl^- is less stable than Cl^{11-}

☐ I do not know

Why did you think this was the answer?

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18. Tick ✓ one of the four statements:

☐ Cl^{11-} is more stable than Cl

☐ Cl^{11-} and Cl are equally stable

☐ Cl^{11-} is less stable than Cl

☐ I do not know

Why did you think this was the answer?

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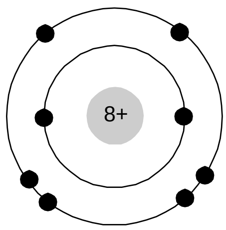
Chemical stability (7)

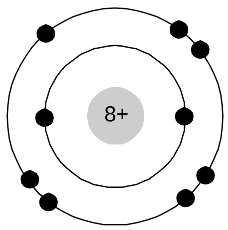
The diagrams below represent three chemical species:

O



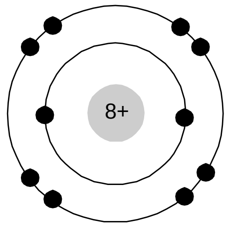
Oxygen atom with electronic configuration of 2.6

O^-



Oxygen ion with electronic configuration of 2.7

O^{2-}



Oxide ion with electronic configuration of 2.8

19. Tick ✓ one of the four statements:

- ☐ O is more stable than O^-
- ☐ O and O^- are equally stable
- ☐ O is less stable than O^-
- ☐ I do not know

Why did you think this was the answer?

20. Tick ✓ one of the four statements:

- ☐ O^- is more stable than O^{2-}
- ☐ O^- and O^{2-} are equally stable
- ☐ O^- is less stable than O^{2-}
- ☐ I do not know

Why did you think this was the answer?

21. Tick ✓ one of the four statements:

- ☐ O^{2-} is more stable than O
- ☐ O^{2-} and O are equally stable
- ☐ O^{2-} is less stable than O
- ☐ I do not know

Why did you think this was the answer?

Activity 7: Rusting - Lateral Thinking Exercise

Corrosion of iron is a major economic problem worldwide. The rust forms on the surface of iron in the presence of moisture and oxygen. There are many experiments available whereby students can investigate what conditions are necessary for rusting to occur. The rust formed is a form of iron(III)oxide.

One of the methods of protecting against rusting is related to the periodic table. If the iron is connected to a metal of higher reactivity, the more reactive metal will corrode first. A familiar example of this is the galvanising of iron with zinc, a more reactive metal. The zinc then 'protects' the iron by becoming corroded first.

The web link below will launch a higher order lateral and creative thinking exercise based around corrosion on a sunken ship and specifically the *Titanic*. The exercise develops questioning skills and lateral thinking as well as the use of concept cartoons as a way of presenting the results of a discussion. The planning of an investigation is also a key aspect of this resource including such areas as 'fair testing'.

http://www.rsc.org/Education/Teachers/Resources/Books/GiftedandTalented.asp

Worksheets for 11-14 year olds

Resource 7 in the listing. This resource includes a powerpoint presentations, worksheets, handouts and experimentation.

Activity 8: Diagnostic test

Part 1

1 Which is more reactive, potassium or magnesium, and why? (7)
2 Which is more reactive, sulfur or fluorine, and why? (7)
3 Why might you expect sulfur and bromine to have similar reactivities? (7)
4 Why can aluminium not be extracted from its oxide ore using coke as a reducing agent? (3)
5 Why does copper not react with dilute hydrochloric acid? (3)

Part 2

1 Complete the following word equations by either writing the names of the products, or 'no reaction' if you think there will be no reaction:

a) Magnesium + copper sulfate →
b) Lead + zinc nitrate →

c) Sodium hydroxide + sodium nitrate →
d) Potassium hydroxide + zinc sulfate →
e) Aluminium + iron oxide →
f) Zinc + sulfuric acid →
g) Copper + hydrochloric acid v
h) Sodium hydroxide + nitric acid →

i) Calcium carbonate + hydrochloric acid →

j) Magnesium + water →

Answers

Part 1

1	Potassium is more reactive as it only has to lose one electron – Mg has to lose two. It is more difficult to lose two than to lose one – K's outermost electron is further from the nucleus than Mg's is – less attraction to nucleus – easier to lose.	**(7)**
2	Fluorine is more reactive as it only has to gain one electron – S has to gain two. It is more difficult to gain two than to *gain one* – F's outermost electron is closer to the nucleus than S's is – more attraction to nucleus – easier to gain.	**(7)**
3	S is the third period – Br is the fourth – as you go from the third to the fourth period on right-hand side of the table, reactivity decreases. Br is on the extreme RHS – S is further to the left – as you go from left to right on the RHS of the table, reactivity increases. Thus, reactivities are roughly the same.	**(7)**
4	Coke is carbon. Aluminium is more reactive than carbon, so carbon cannot displace aluminium from its ore.	**(3)**
5	All acids contain hydrogen. Copper is less reactive than hydrogen, so copper cannot displace hydrogen from dilute acids.	**(3)**

Part 2

1	a)	magnesium sulfate (1) + copper (1)	**(2)**
	b)	no reaction	**(2)**
	c)	no reaction	**(2)**
	d)	potassium sulfate (1) + zinc hydroxide (1)	**(2)**
	e)	aluminium oxide (1) + iron (1)	**(2)**
	f)	zinc sulfate (1) + hydrogen (1)	**(2)**
	g)	no reaction	**(2)**
	h)	sodium nitrate (1) + water (1)	**(2)**
	i)	calcium chloride (1) + water (1) + carbon dioxide (1)	**(3)**
	j)	magnesium oxide/hydroxide (1) + hydrogen (1)	**(2)**

Glossary

Below are some of the terms used within this document:

Term	Definition
corrosion	Corrosion is the disintegration of a material into its constituent components
displacement	A more reactive substance replaces a less reactive one from a compound
electron	Stable elementary particle with a negative charge
element	An type of atom characterised by the number of protons in the nucleus
ion	An atom that has either lost or gained electrons forming a charged particle
oxidation	The loss of electrons
reduction	The gain of electrons
thermal dissociation	A reaction that cannot be reversed by altering the temperature.

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