3.1 Properties of metals

More than 80% of the elements in the periodic table are metals.

Metals have been important to human beings since early times. The development of civilisation can be measured by the way we have used metals. The Copper Age (5000–3000 BCE) was followed by the Bronze Age (3000–1000 BCE) and the Iron Age (from 1000 BCE).

Gold, silver and copper can be found on Earth in an almost pure form. These metals were employed by prehistoric humans to make ornaments, tools and weapons. As humans' knowledge of metallurgy (the science of modifying metals) developed, metals have played a central role in fields as diverse as construction, agriculture, art, medicine and transport.

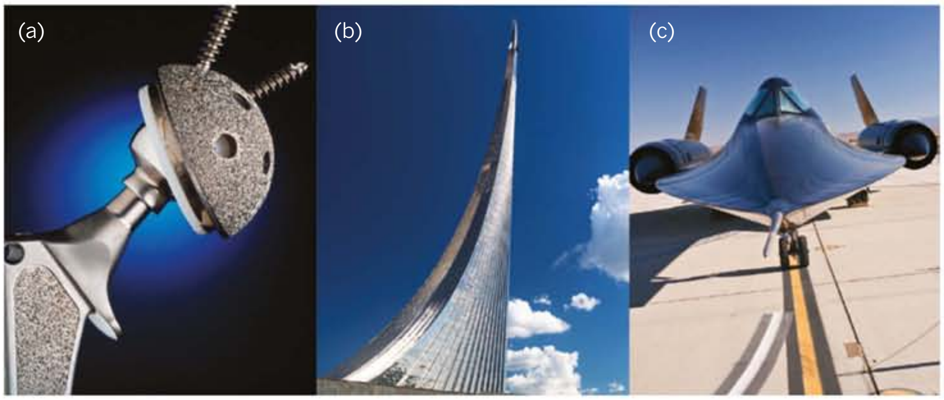


FIGURE 3.1.1 Titanium has many uses: (a) a replacement hip socket, (b) the spectacular curved space museum building in Moscow and (c) the SR-71 Blackbird reconnaissance aircraft. The SR-71 aircraft is the fastest manned aircraft that uses oxygen directly from the atmosphere.

The diverse properties of different metals make them suitable for many purposes. Table 3.1.1 shows the uses of some metals. For example, titanium (Figure 3.1.1) is a very strong, relatively unreactive metal with a low density that is close to that of bone. Consequently, it is used in surgical implants that can last up to 20 years with little effect on the body. Titanium is also used in the aerospace industry, in art and architecture, and in sporting products such as golf clubs.

TABLE 3.1.1 Properties and uses of some metals

Metal	Properties	Uses
Iron	Soft, malleable, magnetic, good thermal and electrical conductor, fairly reactive, readily forms alloys	Can corrode and is usually converted to more stable steel, which is used in buildings and bridges, automobiles, machinery and appliances
Aluminium	Low density, relatively soft when pure, excellent thermal and electrical conductor, malleable and ductile, good reflector of heat and light, readily forms alloys	Saucepans, frying pans, drink cans, cooking foil, food packaging, roofing, window frames, appliance trim, decorative furniture, electrical cables, aircraft and boat construction
Titanium	Very strong, high melting point, low density, low reactivity, readily forms alloys	Medical devices within the body, wheelchairs, computer cases; lightweight alloys are used in high-temperature environments such as spacecraft and aircraft
Gold	Shiny gold appearance, excellent thermal and electrical conductor, unreactive, readily forms alloys	Electrical connections, jewellery, monetary standard, dentistry

In this section, you will examine the properties of metals. Then, in the next section, you will learn about the bonding model that chemists have developed to explain these properties. This model has helped chemists and materials engineers to understand why metals behave the way they do and how metals can be modified to create useful new materials.

PROPERTIES OF METALS

Table 3.1.2 gives the properties of some metals and non-metals. Despite the different properties of metals, most metals:

- exhibit a range of melting points and relatively high boiling points
- are good **conductors** of electricity
- are good conductors of heat
- generally have high **densities**.

TABLE 3.1.2 Properties of some metallic and non-metallic elements

Element	Melting point (°C)	Boiling point (°C)	Electrical conductivity ($MS\ m^{-1}$)*	Thermal conductivity ($J\ s^{-1}\ m^{-1}\ K^{-1}$)†	Density ($g\ mL^{-1}$)
Metals					
Gold	1063	2970	45	310	19.3
Iron	1540	3000	9.6	78	7.86
Mercury	–39	357	1	8.4	13.5
Potassium	64	760	14	100	0.86
Silver	961	2210	60	418	10.5
Sodium	98	892	21	135	0.97
Non-metals					
Carbon (diamond)	3550	‡	10^{-17}	—	3.51
Oxygen	–219	183	—	0.026	1.15 (liquid)

*$MS\ m^{-1}$ = megasiemens per metre.
†Thermal conductivity measures the conductance of heat.
‡Diamond sublimes (changes straight from a solid to a gas) when heated.

Not all metals have all of these properties. Mercury is a liquid at room temperature—it has an unusually low melting point. The group 1 elements (the **alkali metals**) have some properties that make them different from most other metals. They are all soft enough to be cut with a knife and they react vigorously with water to produce hydrogen gas. Both mercury and the group 1 elements exhibit most of the other properties listed above and are classified as metals.

Metals also generally have the following characteristics in common. They:

- are **malleable**—they can be shaped by beating or rolling
- are **ductile**—they can be drawn into a wire
- are lustrous or reflective when freshly cut or polished
- are often hard, with high **tensile strength**
- have low **ionisation energies** and electronegativities.

These properties can allow different metals to be used together in order to solve many engineering problems. The power transmission tower in Figure 3.1.2 is made of a few metals to take advantage of their different properties.

FIGURE 3.1.2 This power transmission tower relies on the strength of iron in steel for its structural integrity. The electricity cables are made from aluminium, utilising its ductility and electrical conductivity.

Generally, metals are shaped for use in different applications by hammering, exploiting their malleability. Some metals, such as gold, copper and aluminium, are very malleable at room temperature. Other metals, such as iron, must be heated before they can be shaped.

Generally, metals are good conductors of electricity and heat. They are malleable, ductile, and have high tensile strength and low ionisation energies.

Most metals are similar in appearance, being lustrous (reflective) and silvery-grey. Gold and copper are notable exceptions. Gold is a yellow coloured metal; copper is reddish.

FORMING METAL IONS

In Chapter 1, you saw that there is a general tendency for atoms to combine so that they have eight electrons in the outer shell (the **octet rule**).

Metallic elements are found on the left-hand side of the periodic table. The atoms of metals are generally larger than the atoms of non-metallic elements within a period and the **core charge** of their atoms is lower. It takes less energy to remove electrons from an outer shell when an atom is large, so the ionisation energy of metals is usually lower than for non-metals in the same period. As a consequence, metal atoms tend to lose their outer-shell electrons to form positive ions, called **cations**.

Atoms of simple metals have one, two or three electrons in their outer shell. The cations that are formed when these metal atoms lose these valence electrons have a stable noble gas electronic configuration, with eight electrons in their outer shell.

Worked example 3.1.1

DETERMINING CHARGES

Determine the charge of a calcium cation.

Thinking	Working
Unreacted calcium atoms have the same number of protons and electrons.	Atomic number (Z) of calcium is 20: number of protons is 20, number of electrons is 20
The electrons in an atom are in shells.	Shell configuration of calcium: 2,8,8,2
Only the outer-shell electrons will be lost.	Outer shell contains two electrons, 20 – 2 = 18 electrons remaining
Cation charge = number of protons – number of electrons	Cation charge = 20 – 18 = +2

Worked example: Try yourself 3.1.1

DETERMINING CHARGES

Determine the charge of an aluminium cation.

CHEMFILE

Beryllium

Beryllium is the fourth element in the periodic table. It is one of the lightest metals, its density being two-thirds the density of aluminium. Beryllium is non-magnetic and has six times the stiffness of steel. The Space Shuttle and the Spitzer Space Telescope both use beryllium due to its strength and light weight. NASA's next-generation James Webb Space Telescope shown in Figure 3.1.3, scheduled for launch in 2018, will depend on a 6.5-metre mirror constructed using beryllium to see objects 200 times fainter than those previously visible.

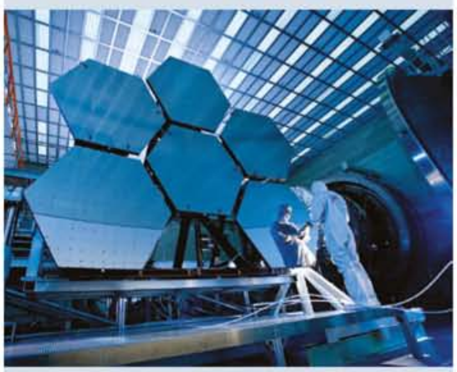


FIGURE 3.1.3 Some of the 18 mirror segments of the James Webb Space Telescope. The mirrors are supported by beryllium ribs that maintain the mirror's shape under extreme conditions.

TRANSITION METALS

Between group 2 and group 13 in the periodic table is a block of elements known as the **transition metals**. These elements generally have unfilled d-subshells and are often referred to as the d-block. (See the periodic table at the end of the book.) They include metals such as iron and nickel that are used to build cities, bridges, cars and railway lines, and precious metals such as silver and gold that have ornamental and economic uses. Most transition metals are silver-coloured and are similar in appearance, as can be seen in Figure 3.1.4.

FIGURE 3.1.4 The first row of transition metals.

The transition metals are very important to Australian industry. All of the metals in the first row are found in Australia and many are being mined today.

Iron is by far the most important metal to us. Nearly ten times more iron is mined than all other metals combined. Iron obtained directly from a blast furnace is relatively **brittle** and corrodes easily. Carbon and other transition metals are combined with iron to produce mixtures or **alloys**, called **steel**. Steel has more desirable characteristics than pure iron. (Alloys are covered in more detail in section 3.5.)

Copper is one of the few transition metals that is mainly used in its pure form. It is highly electrical conductive and so it is used for most of the millions of kilometres of electrical wires that enable the transmission of electric energy for heating, lighting, telephones, radio and television.

Transition metals are not only important for industry, they are also important for life. All the transition metals in the first row, except scandium and titanium, are essential for animal life. Your body relies on the presence of trace elements to carry out certain biochemical reactions. For example, chromium, which you get from meat and bread, assists in the production of energy from glucose.

Properties of transition metals

Compared to the main group metals, transition metals have the following properties.

- They tend to be harder.
- They have higher densities.
- They have higher melting points.
- Some of them have strong magnetic properties.

The hardness, higher densities and higher melting points are due to the atoms of transition metals generally being a smaller size due to their greater core charge. This allows them to pack together more tightly with stronger bonds.

The high tensile strength of transition metals makes them suitable for use in the construction of buildings, cars, bridges and numerous other objects.

Transition metal compounds

Transition metal compounds display a wide range of different colours. They are extensively used as pigments in paints, and to colour glass, ceramics and enamel. In Figure 3.1.5 on page 58, the colours used by the artists are caused by the different transition metals present and the colours are still as vivid today as when they were painted.

FIGURE 3.1.5 (a) Arthur Streeton's 'Golden summer, Eaglemont', 1889. (b) An Australian Aboriginal abstract painting.

Ochre is a type of hard clay that contains iron oxides and hydroxides, which can be found naturally in many colours, including red, pink, white and yellow. Ground into a powder and mixed with liquids, ochre forms a paste that has been used for millennia by Aboriginal and Torres Strait Islander people in Australia for body decoration, cave painting, bark painting and other artwork.

The colours of many gemstones are also due to the presence of transition metals. For example, sapphires (Figure 3.1.6) contain traces of titanium and iron in a crystal **lattice** of aluminium oxide.

The colours arise when electrons within the metal ions in the compounds absorb light of particular wavelengths and move to higher energy levels. Absorbance of light with some wavelengths and transmission of light with other wavelengths results in the compounds appearing coloured. By contrast, compounds of group 1 and group 2 metals are usually colourless.

FIGURE 3.1.6 Blue sapphires get their colour from impurities of titanium and iron.

3.1 Review

SUMMARY

- Metals have the following characteristic properties:
  - high boiling points
  - good conductors of electricity in solid and liquid states
  - malleable and ductile
  - high densities
  - good conductors of heat
  - lustrous
  - low electronegativities
  - low ionisation energies
  - react by losing electrons.
- The main differences between the properties of main group and transition metals are:
  - transition metals are harder
  - transition metals are more dense
  - transition metals have higher melting points
  - some transition metals have strong magnetic properties
  - transition metal compounds tend to be brightly coloured.

KEY QUESTIONS

1 Determine the charge of the cations formed from the following metals if they lost all of their outer-shell electrons.
  a Li
  b Mg
  c Ga
  d Ba

2 **a** Potassium is classed as a metal. Which of its properties are similar to those of the metal gold? In what ways is it different?
  b Identify another element in Table 3.1.2 on page 55 that has similar properties to potassium.
  c Identify another metal in Table 3.1.2 on page 55 that has similar properties to gold.
  d Where are these four metals in the periodic table?

3 **a** Which metals would you select if you wanted a good electrical conductor?
  b What other factors might influence your choice?

4 Sodium and iron have very different physical properties. Explain why this is so based on where these metals are found in the periodic table.

5 Suggest some properties not included in Table 3.1.2 on page 55 that you would need to consider before choosing between aluminium and iron for building a bridge.

3.2 Metallic bonding

In this section, you will learn how the properties of metals can be explained in terms of the structure of the particles in metals. You will also learn about the bonding model that chemists have developed to explain these properties. The metallic bonding model has helped chemists and materials engineers to understand why metals behave as they do and how metals can be modified to create useful new materials.

CONNECTING PROPERTIES AND STRUCTURE

The properties of metals are listed in Table 3.2.1. Each of these properties gives some information about the structure and bonding of particles in metals.

TABLE 3.2.1 The physical properties of metals and resulting conclusions about metal structure and bonding

Property	What this tells us about structure
Metals are usually hard and tend to have high boiling points.	The forces between the particles must be strong.
Metals conduct electricity in the solid state and in the molten liquid state.	Metals have charged particles that are free to move.
Metals are malleable and ductile.	The attractive forces between the particles must be stronger than the repulsive forces between the particles when the layers of particles are moved.
Metals generally have high densities.	The particles are closely packed in a metal.
Metals are good conductors of heat.	There must be a way of quickly transferring energy throughout a metal object.
Metals are lustrous or reflective.	Free electrons are present, so metals can reflect light and appear shiny.
Metals tend to react by losing electrons.	Electrons must be relatively easily removed from metal atoms.

Chemists have developed a model for the structure of metals to explain all the properties that have been mentioned so far. You can deduce from the information in Table 3.2.1 that the metallic bonding model must include:

- charged particles that are free to move and conduct electricity
- strong forces of attraction between atoms throughout the metal structure
- some electrons that are relatively easily removed.

METALLIC BONDING MODEL

Electrons are the particles that enable metals to conduct electricity. They are able to move between the lattice (tightly packed arrangement) of metal atoms. Negatively charged electrons can be lost from the outer shell of metal atoms, forming positive ions (cations). As shown in Figure 3.2.1, the freed electrons **delocalise** (spread through a large area) to form a 'sea' of electrons throughout the entire metal structure due to the strong attraction to the metal cations.

Chemists believe that, in a solid sample of a metal:

- positive ions are arranged in a closely packed structure. This structure is described as a regular, three-dimensional network of positive ions. The cations occupy fixed positions in the lattice
- negatively charged electrons move freely throughout the lattice. These electrons are called **delocalised electrons** because they belong to the lattice as a whole, rather than staying in the shell of a particular atom
- the delocalised electrons come from the outer shells of the atoms. Inner-shell electrons are not free to move throughout the lattice and remain firmly bonded to individual cations

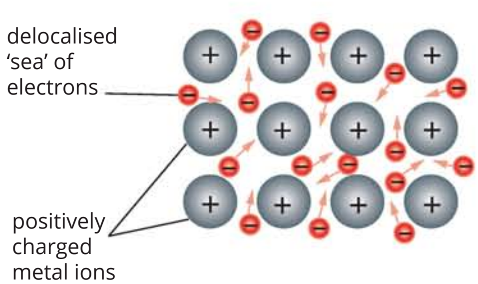


FIGURE 3.2.1 The metallic bonding model. Positive metal cations are surrounded by a mobile sea of delocalised electrons. This diagram shows just one layer of metal ions.

- the positive cations are held in the lattice by the electrostatic force of attraction between these cations and the delocalised electrons. This attraction extends throughout the lattice and is called **metallic bonding**.

Together, these ideas make up the metallic bonding model. An example of how a metal, such as sodium, could be represented using this model is shown in Figure 3.2.2.

> In the metallic bonding model, positive metal cations are surrounded by a sea of delocalised electrons.

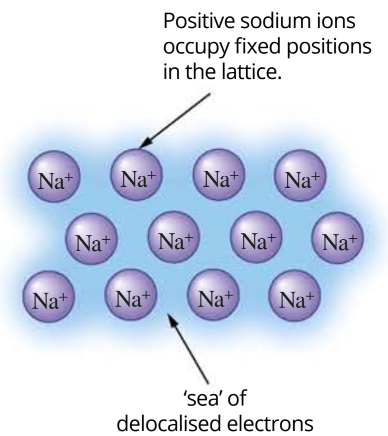
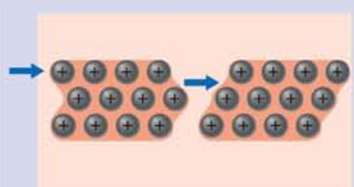


FIGURE 3.2.2 A representation of a sodium metal lattice. Each sodium atom loses its one valence electron. This electron is shared with all atoms in the lattice to form a sea of delocalised electrons.

EXPLAINING THE PROPERTIES OF METALS

Table 3.2.2 shows how the metallic bonding model is consistent with the relatively high boiling point, electrical conductivity, malleability and ductility of metals.

TABLE 3.2.2 Physical properties of metals and explanations from metallic bonding model

Property	Explanation	
Metals are hard and have relatively high boiling points.	Strong electrostatic forces of attraction between positive metal ions and the sea of delocalised electrons holds the metallic lattice together.	
Metals are good conductors of electricity.	Free-moving delocalised electrons will move towards a positive electrode and away from a negative electrode in an electric circuit.	
Metals are malleable and ductile.	When a force causes metal ions to move past each other, layers of ions are still held together by the delocalised electrons between them.	

Other properties of metals

Metals generally have a high density. The cations in a metal lattice are closely packed. The density of a metal depends on the mass of the metal ions, their radius and the way in which they are packed in the lattice.

Metals are good conductors of heat. When the delocalised electrons bump into each other and into the metal ions, they transfer energy to their neighbour. Heating a metal gives the ions and electrons more energy and they vibrate more rapidly. The electrons, being free to move, transmit this energy rapidly throughout the lattice.

Metals are lustrous. Because of the presence of free electrons in the lattice, metals reflect light of all wavelengths and appear shiny.

Metals tend to react by losing electrons. The delocalised electrons in metals may participate in reactions anywhere on the metal's surface. The reactivity of a metal depends on how easily electrons can be removed from its atoms. This is covered in more detail in section 3.3.

Limitations of the metallic bonding model

Although this model of metallic bonding explains many properties of metals, some cannot be explained so simply. These include the:

- range of melting points, hardness and densities of different metals

- differences in electrical conductivities of metals
- magnetic nature of metals such as cobalt, iron and nickel.

In order to deal with such questions, you would need a more complex model of metallic bonding, which is beyond the scope of this book.

Worked example 3.2.1

ELECTRONIC CONFIGURATION OF ALUMINIUM

With reference to the electronic configuration of aluminium, explain why solid aluminium can conduct electricity.

Thinking	Working
Using the atomic number of the element, determine the electronic configuration of its atoms. (You may need to refer to a periodic table.)	Al has an atomic number of 13. This means a neutral atom of aluminium has 13 electrons. The electronic configuration is $1s^22s^22p^63s^23p^1$.
From the electronic configuration, find how many outer-shell electrons are lost to form cations that have a stable, noble gas electronic configuration. These electrons become delocalised.	Al has 3 electrons in its outer shell (the $3s^23p^1$ electrons). Al atoms will tend to lose these 3 valence electrons to form a cation with a charge of 3+. The outer-shell electrons become delocalised and form the sea of delocalised electrons within the metal lattice.
An electric current occurs when there are free-moving charged particles.	If the Al is part of an electric circuit, the delocalised electrons are able to move through the lattice towards a positively charged electrode.

Worked example: Try yourself 3.2.1

ELECTRONIC CONFIGURATION OF MAGNESIUM

With reference to the electronic configuration of magnesium, explain why solid magnesium can conduct electricity.

3.2 Review

SUMMARY

- Metallic bonding is the electrostatic force of attraction between a lattice of positive ions and delocalised valence electrons. The lattice of cations is surrounded by a sea of delocalised electrons.
- The metallic bonding model can be used to explain the properties of metals, including their malleability, thermal conductivity, generally high melting point and electrical conductivity.

KEY QUESTIONS

1 The properties of calcium mean that it is classed as a metal.
  a Draw a diagram to represent a calcium metal lattice.
  b Describe the forces that hold this lattice together.

2 Barium is an element in group 2 of the periodic table. It has a melting point of 850°C and conducts electricity in the solid state. Describe how the properties of barium can be explained in terms of its bonding and structure.

3 Graphite is a non-metallic substance that can be lustrous and conducts electricity and heat. It is not malleable, but breaks if a force is applied.
  a What properties does graphite share with metals?
  b What inferences can you make about the structure of graphite given it shares these properties with metals?

3.3 Reactivity of metals

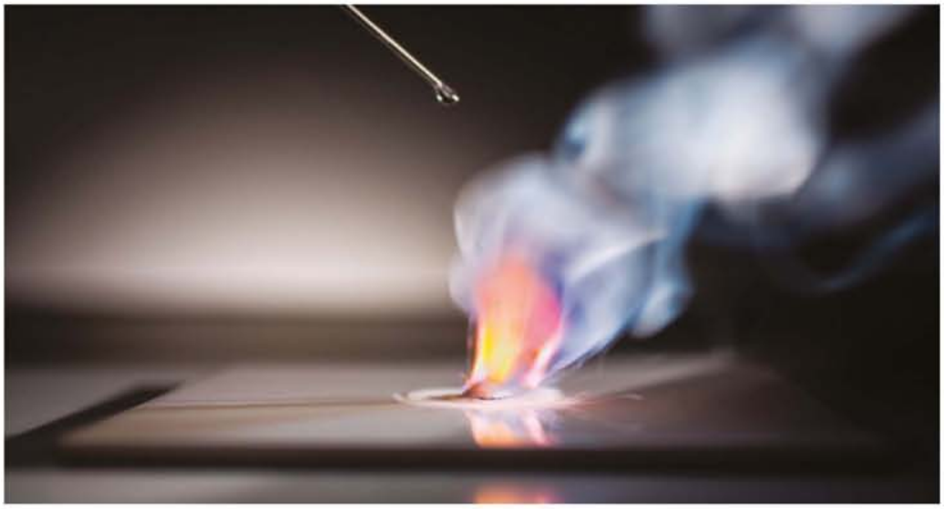
In the previous section, you learned that metals have many common properties. Metallic elements can also have very different properties. These include their **reactivity** with water, acids and oxygen. Some metals are extremely reactive and others are much less so.

This section will look at how the reactivity of different metals can be determined experimentally and what are some of the periodic patterns that exist.

DETERMINING THE REACTIVITY OF METALS

Reactivity with water

The way metals react with water can indicate their relative reactivity.



potassium + water → potassium hydroxide + hydrogen gas

FIGURE 3.3.1 When water is dropped onto metallic potassium, hydrogen gas is produced.

CHEMFILE

Reactivity of group 1 metals

Group 1 metals are so reactive that they must be handled with great care. They need to be stored under oil to prevent the metal coming into contact with moisture in the atmosphere.

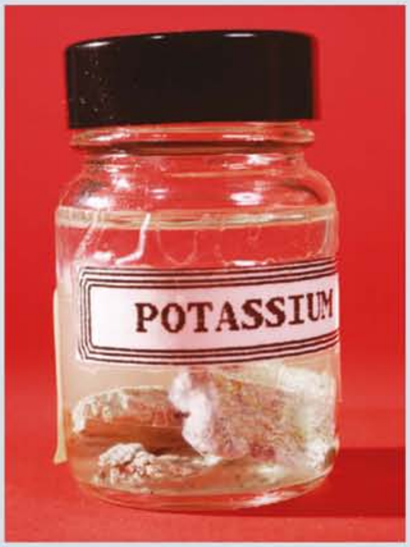


FIGURE 3.3.2 Potassium metal is stored under oil to prevent contact with moisture.

Figure 3.3.1 shows the reaction of potassium, a group 1 metal, with water. Enough heat is generated to instantly melt the potassium and ignite the hydrogen. The vigour of the reaction is an indication of the reactivity of the metal. Potassium has high reactivity with water, which is characteristic of the group 1 metals.

Table 3.3.1 describes the reaction of some group 1 and group 2 metals with water. In each case, a reaction results in the formation of hydrogen gas.

TABLE 3.3.1 Reaction of selected group 1 and 2 metals with water

Period	Group	Element	Reaction with water
3	1	Sodium	Reacts vigorously, producing enough energy to melt the sodium, which fizzes and skates on the water surface
4	1	Potassium	Reacts violently, making crackling sounds as the heat evolved ignites the hydrogen produced by the reaction
5	1	Rubidium	Explodes violently on contact with water
3	2	Magnesium	Will not react with water at room temperature but will react with steam
4	2	Calcium	Reacts slowly with water at room temperature

From these and other experimental observations we can say that, in general:

- metals in group 1 of the periodic table (i.e. Na, K and Rb) are more reactive in water than those in group 2 (i.e. Mg and Ca)
- going down a group, the reactivity of the metal in water increases.

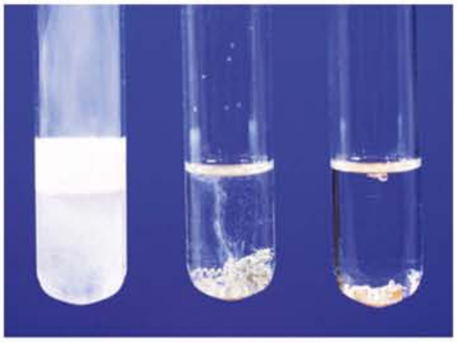
Transition metals

Transition metals are generally less reactive with water than group 1 and 2 metals are. For example, iron reacts fairly slowly with water. Gold and platinum are essentially unreactive.

Reactivity with acids

The reactivity of different metals with acids follows the same general patterns as the reactivity of metals with water. Metals are normally more reactive with acids than with water. More metals react with acids and the reactions tend to be more energetic.

Metals can be placed in an order of their relative reactivity. In Figure 3.3.3, the reactions of magnesium, iron and copper with an acid are shown. The large amount of bubbling and the mist produced show that magnesium is the most reactive metal of the three, whereas copper is the least.



magnesium iron copper

FIGURE 3.3.3 Metals reacting with an equal amount of dilute acid. From left to right: magnesium ribbon, iron filings and copper turnings.

Reactivity with oxygen

Many metals also react with oxygen. The group 1 metals all react rapidly with oxygen. Figure 3.3.4 shows sodium metal burning in a container of pure oxygen. The metal atoms and oxygen molecules rearrange to form a new compound, sodium oxide. The word equation for this reaction is:

$$\text{sodium} + \text{oxygen} \rightarrow \text{sodium oxide}$$

FIGURE 3.3.4 Sodium burning in pure oxygen.

The group 2 metals also react with oxygen to form oxides (compounds containing the O^{2-} **anion**), although not as rapidly as group 1 metals. Heat is usually required to start the reaction.

While the transition metals are less reactive with oxygen than the metals in groups 1 and 2, their reactions are also important. Iron forms rust (iron oxide) when exposed to oxygen and water over a period of time. Many transition metals needed by society cannot be found in nature as a pure element but often exist as oxides.

Iron, copper, titanium and aluminium are all mined as the oxides and must be processed to obtain the finished metal. Figure 3.3.5 shows a cluster of crystals of rutile, the oxide from which titanium is extracted. The production of iron from iron oxide is covered in detail in section 3.4.

FIGURE 3.3.5 Rutile is a mineral that contains titanium dioxide.

Gold and platinum, which are much less reactive than most other metals, are found in the Earth's crust in their pure form. Gold is often found in rock formations called seams alongside quartz, as shown in Figure 3.3.6.

FIGURE 3.3.6 Gold, an unreactive metal, exists in the Earth's crust in its metallic elemental form.