Unit 11 Reactivity of metals

Comparing the reactivity of metals

In unit 4, you learned that different extraction methods are used in the extractions of metals. Some metals, like silver and gold can exist as free elements on earth and can be extracted by physical methods. However, other metals, such as iron has to be extracted by heating its oxide with carbon. Why does iron not exist as free elements on earth? The reason is iron is more reactive than silver and gold. It would combine with other elements to form compounds on earth. It shows that different metals would have different reactivity.

To compare the reactivity of metals, we usually consider the following three aspects:

1. **The temperature at which the reaction starts**
   Many reactions start only when heat is applied. The more reactive the metal is, the lower the temperature required.

2. **The rate of reaction**
   A more reactive metal reacts faster than a less reactive one.

3. **The amount of heat released during reaction**
   Generally, the more reactive the metal is, the more heat will be released during reaction.

In this unit, we will look at the reactions of metals with oxygen, water / steam, dilute acids; the displacements reactions of metals and the reduction of metal oxides in order to investigate the difference in reactivity of different metals.

The diagram below summarizes the typical reactions of metals with oxygen, water / steam and dilute acids.
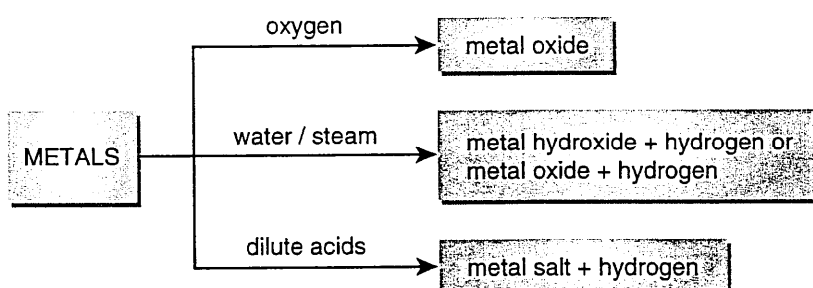
The reactions of metals with oxygen in the air

Most metals react with oxygen in the air to produce metal oxide. The reactions usually required heat to start. The following word equation can describe the result if there is a reaction:

$$\text{metal + oxygen} \longrightarrow \text{metal oxide}$$
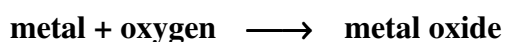
You can see that the metal only reacts with oxygen in the air. Nitrogen in the air is too unreactive and generally does not have reactions with most metals.

The table on the next page lists the observations when different metals are gently heated. It also lists the word equations if chemical reactions do occur. You can see that some metals react more readily than the others.

Metal	Heat needed to start reaction	Observation	Word equation for the reaction
Potassium (K)	gentle	potassium burns vigorously with a *lilac* flame; a white smoke forms	potassium + oxygen ⟶ potassium oxide (K_2O)
Sodium (Na)	gentle	sodium burns vigorously with a golden yellow flame; a white smoke forms	sodium + oxygen ⟶ sodium oxide (Na_2O)
Calcium (Ca)	strong	calcium burns quite vigorously with a brick-red flame; a white powder forms	calcium + oxygen ⟶ calcium oxide (CaO)
Magnesium (Mg)	strong	magnesium burns with a very bright light; a white powder forms	magnesium + oxygen ⟶ magnesium oxide (MgO)
Aluminium (Al)	strong	a white powder forms	aluminium + oxygen ⟶ aluminium oxide (Al_2O_3)
Zinc (Zn)	strong	a powder (yellow when hot but white when cold) forms	zinc + oxygen ⟶ zinc oxide (ZnO)
Iron (Fe)	strong	iron powder burns with sparks; a black solid forms	iron + oxygen ⟶ iron(II, III) oxide* (Fe_3O_4 or $FeO \cdot Fe_2O_3$)
Lead (Pb)	strong	a powder (orange when hot but yellow when cold) forms on the surface	lead + oxygen ⟶ lead(II) oxide (PbO)
Copper (Cu)	very strong	a black powder forms on the surface	copper + oxygen ⟶ copper(II) oxide (CuO)
Mercury (Hg)	very strong	a red powder forms on the surface	mercury + oxygen ⟶ mercury(II) oxide (HgO)
Silver (Ag)	—	no observable change	—
Platinum (Pt)	—	no observable change	—
Gold (Au)	—	no observable change	—

Order of reactivity of different metals with oxygen in the air
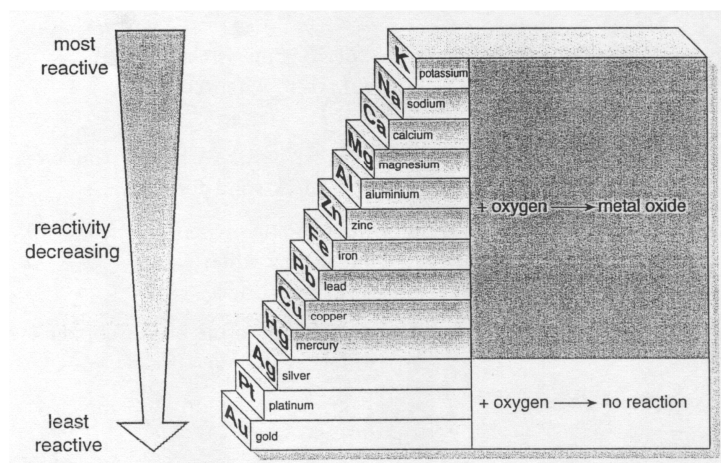
Many metals look dull. It is because these metals react with oxygen in the air to form an oxide layer. When these metals are freshly cut, they appear shiny. Gold is the least reactive metal. It does not react with other substances in the environment. Therefore, it always appears shiny.

Potassium and sodium are very reactive. They must be stored in paraffin oil to prevent their reactions with oxygen in the air on storing.

**The reactions of metals with water (or steam)**

Some metals react cold water. Some metals can only react with hot water or steam. Many metals cannot react with water or even steam.

Reactions between water and potassium / sodium / calcium

Potassium, sodium and calcium react with cold water, producing metal hydroxides and hydrogen gas. metal + water $\longrightarrow$ metal hydroxide + hydrogen

The diagram on the right hand side shows the experimental set-up for the reaction between calcium and water.
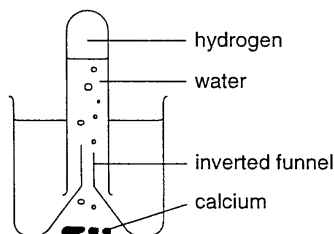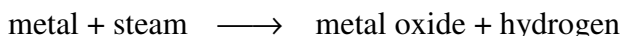

Reaction between steam and magnesium / aluminium* / zinc / iron

Magnesium reacts very slowly with cold water producing only a few bubbles.

Aluminium, zinc and iron do not react with cold water.

However, these metals (except aluminium) react readily with steam. Metal oxides and hydrogen are produced during the reaction.

metal + steam $\longrightarrow$ metal oxide + hydrogen

* Aluminium is often covered with a thin layer of oxide and this prevents its reaction with steam. Aluminium can react with steam if the oxide layer is removed. The oxide can be removed by dipping the metal into a mixture of mercury(II) chloride solution and dilute hydrochloric acid.

The table below lists the observations when metals react with water / steam. It also lists the word equations if chemical reactions do occur. You can see that some metals react more readily than the others.

Metal	Observation	Word equation for the reaction
Potassium (K)	• potassium melts to form a silvery bead • the bead moves rapidly on the water surface • potassium burns with a lilac flame	potassium + water $\longrightarrow$ potassium hydroxide + hydrogen (KOH)
Sodium (Na)	• sodium melts to form a silvery bead • the bead moves rapidly on the water surface • sodium burns with a golden yellow flame	sodium + water $\longrightarrow$ sodium hydroxide + hydrogen (NaOH)
Calcium (Ca)	• calcium sinks in water • a steady stream of bubbles forms	calcium + water $\longrightarrow$ calcium hydroxide + hydrogen (Ca(OH)$_2$)
Magnesium (Mg)	• magnesium, aluminium, zinc and iron react readily with steam • the reactivity decreases down the series	magnesium + steam $\longrightarrow$ magnesium oxide + hydrogen (MgO)
Aluminium (Al)		aluminium + steam $\longrightarrow$ aluminium oxide + hydrogen (Al$_2$O$_3$)
Zinc (Zn)		zinc + steam $\longrightarrow$ zinc oxide + hydrogen (ZnO)
Iron (Fe)		iron + steam $\longrightarrow$ iron(II,III) oxide + hydrogen (Fe$_3$O$_4$)

The diagram shows the experimental set up for the reaction between metal and steam
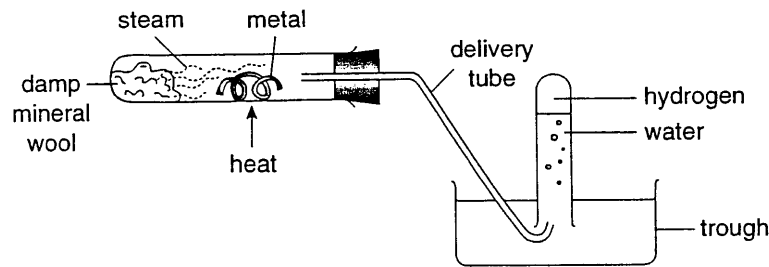
## Order of reactivity of different metals with water / steam
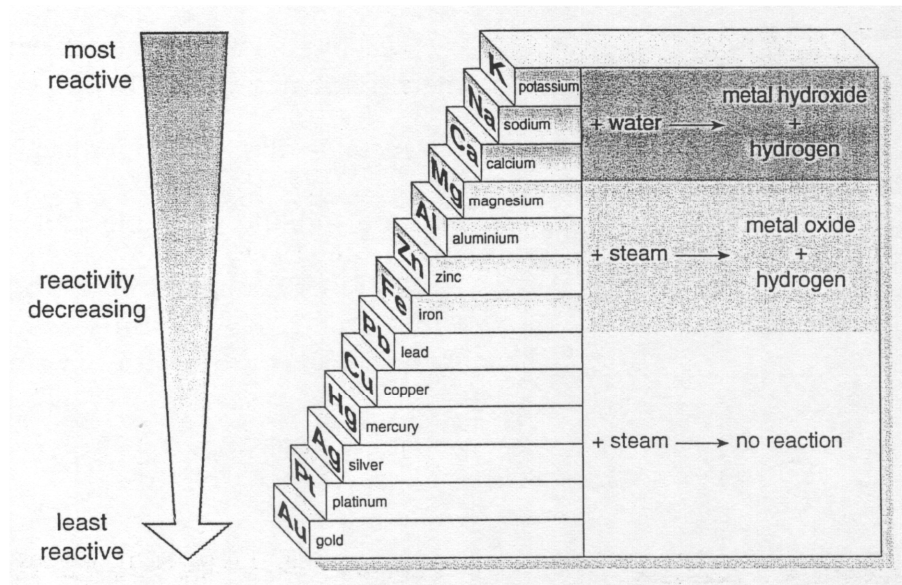
Some metals, such as lead, copper, mercury, silver, platinum and gold do not react with cold water or even steam.

The reactions of metals with dilute acids

Reactive metals react with dilute acids (except dilute nitric acid) to produce salts and hydrogen gas. The kind of salts produced depends on the dilute acid used.

For example, if dilute hydrochloric acid (HCl) is used, a salt called chloride will be produced.

metal + hydrochloric acid $\longrightarrow$ metal chloride + hydrogen
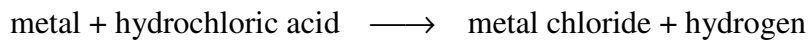
e.g. magnesium + hydrochloric acid $\longrightarrow$ magnesium chloride + hydrogen

Another example: if dilute sulphuric acid (H_2SO_4) is used, a salt called sulphate will be produced.

metal + sulphuric acid $\longrightarrow$ metal sulphate + hydrogen
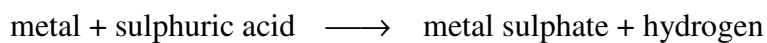
e.g. iron + sulphuric acid $\longrightarrow$ iron(II) sulphate + hydrogen

In all these reactions, the metal will react and dissolve. Colourless bubbles are evolved. The solution will become hot for reactive metals as the reactions would be very vigorous.

Order of reactivity of different metals with dilute acids
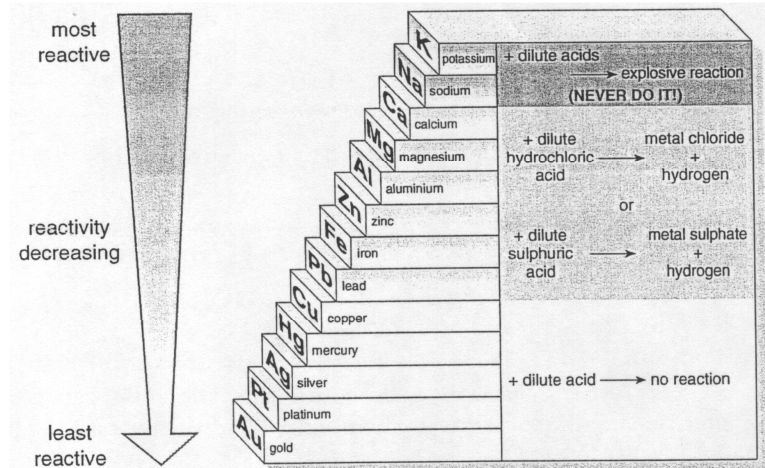
Potassium and sodium **react explosively** with dilute acids. These reactions are **NEVER** carried out in the laboratory.

Some metals, such as copper, mercury, silver, platinum and gold do not react with cold water or even steam.

The metal reactivity series

From the reacts of metals with oxygen, water / steam and dilute acids, it can be seen that metals can be arranged in a list of decreasing reactivity. This order is called the reactivity series. It is shown in the diagram on the right hand side.
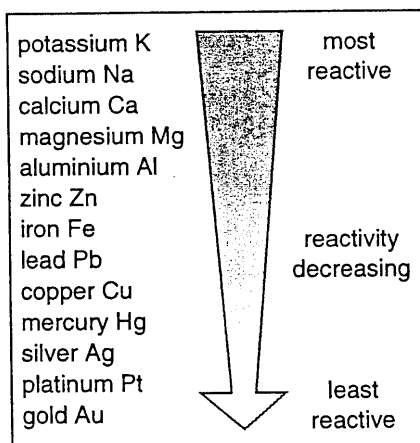
potassium K	most reactive
sodium Na	
calcium Ca	
magnesium Mg	
aluminium Al	
zinc Zn	
iron Fe	reactivity decreasing
lead Pb	
copper Cu	
mercury Hg	
silver Ag	
platinum Pt	least
gold Au	reactive

What is a chemical equation?

You have learnt something about word equation before. It is a simplified way to describe a chemical reaction. For example, when carbon is burnt in the air, it will react with oxygen in the air and carbon dioxide is produced. The reaction can be described by the word equation below.

$$\text{carbon} + \text{oxygen} \longrightarrow \text{carbon dioxide}$$
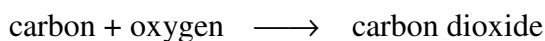
In the above reaction, **carbon and oxygen** are the **reactants** of the reaction. It is written on the left hand side of the word equation.

Carbon dioxide is the **product** of the reaction. It is written on the right hand side of the word equation.

The single arrow '$\longrightarrow$' means is/are changed to.

If the reaction is being investigated at the molecular level, the following changes occur:

| 1 carbon atom (reactant) | 1 oxygen molecule (reactant) | 1 carbon dioxide molecule (product) |
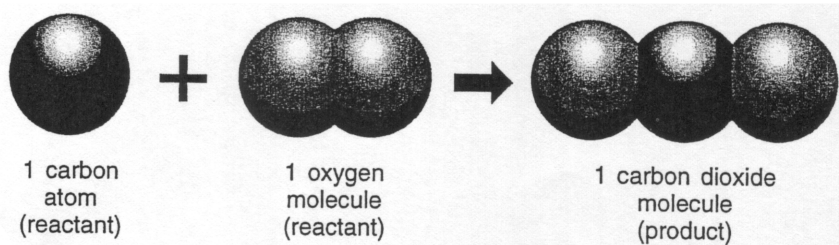
The word equation can be further simplified by using formulae and figures and the reactants and products are adjusted to the correct ratio as the actual reaction carried out. A chemical equation will be produced. The chemical equation for the above reaction is listed below.

$$C + O_2 \longrightarrow CO_2$$
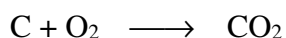
What is a balanced chemical equation?

Consider the reaction between hydrogen and oxygen to form water.

$$\text{hydrogen} + \text{oxygen} \longrightarrow \text{water}$$

If the reaction is being investigated at the molecular level, the following changes occur:

| 2 hydrogen molecules (reactant) | 1 oxygen molecule (reactant) | 2 water molecules (product) |
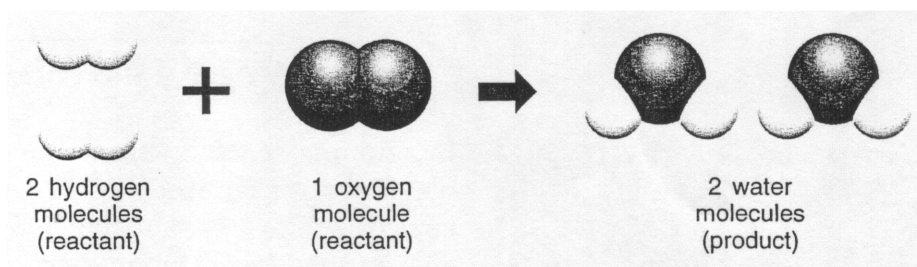
The reaction can be represented by the chemical equation below:

$$2H_2 + O_2 \longrightarrow 2H_2O$$
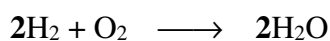
Notice that the **coefficient '2'** has been added in front of H_2 and H_2O.

If we count the total number of atoms in the reactants, it is found that there are altogether $2 \times 2 = 4$ hydrogen atoms and 2 oxygen atoms.

If we count the total number of atoms in the products, it is found that there are altogether $2 \times 2 = 4$ hydrogen atoms and $2 \times 1 = 2$ oxygen atoms.

It can be seen that the total number of each atom on the reactant side is exactly equal to that on the product side. If a chemical equation is written in this state, it is called a balanced chemical equation. Actually, **all chemical equations must be balanced chemical equations**.

Adding more information to chemical equation

The states of reactants may affect the results of a reaction. For example, magnesium does not react readily with liquid water but reacts readily with gaseous steam. Therefore, we might write down the states, using state symbols, of the reactants and the products in the equations.

The table below lists some common state symbols in use and their meanings

State symbol	(s)	(l)	(g)	(aq)
Meaning	Solid	Liquid	Gas	Aqueous solution

The use of state symbols is optional in writing chemical equations.

Obtaining useful information from a chemical equation

Chemical equations give us a lot of useful information. Consider the following chemical equation:

$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(l)$$
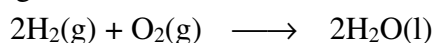
From it, we can obtain the following information:
1. **The reactant(s) involved.**
   There are two reactants, hydrogen and oxygen.
2. **The product(s) formed.** There is only one product water.
3. **The simple ratios of the reactants and products involved in the reaction.**
   e.g. ratio of hydrogen to oxygen = 2 : 1 ;   ratio of hydrogen to water = 2 : 2 or 1 : 1
4. **The states of the substances involved.**
   Hydrogen and oxygen are gases. Water is a liquid.

The single arrow ($\longrightarrow$) in a chemical equation shows that the reaction goes in one way only. All the reactants will be converted to the products if they are mixed in the ratio shown in the equation. This is called '**irreversible**' reaction. Most reactions belong to this type.

However, there can be chemical reactions in which they will actually never complete. Only part of the reactants will be converted to the products and the reactions seem stop. This type of reactions is called '**reversible**' reactions. In writing the equations for this type of reactions, the arrow will be replaced by a double arrow: ($\rightleftharpoons$).

e.g. $3H_2(g) + N_2(g) \rightleftharpoons 2NH_3(g)$

How to write balanced chemical equations

In general, chemical equations may be generated from word equations which are inside our mind.

Rules for writing chemical equations

To write a balanced chemical equation for a reaction, we should follow the steps outlined below. Take the reaction between magnesium and oxygen as an example. Magnesium burns in oxygen to form magnesium oxide.

Step	Example
1 Write down the word equation. On the left, write down the name(s) of the reactant(s). On the right, write down the name(s) of the product(s). Separate the two sides by an arrow ($\longrightarrow$).	magnesium + oxygen $\longrightarrow$ magnesium oxide
2 Write down the chemical formulae of the reactant(s) and product(s).	$Mg + O_2$* $\longrightarrow$ MgO
3 Balance the chemical equation so that the number of each type of atom on one side is equal to that on the other side.	a) There are two oxygen atoms on the reactant side but only one on the product side. To balance the oxygen atom, put the coefficient '2' before MgO*. $Mg + O_2 \longrightarrow 2MgO$ b) Now balance the magnesium atom by putting the coefficient '2' before Mg. $2Mg + O_2 \longrightarrow 2MgO$　　　(Balanced)
4 Write down the state symbol after each substance. (s) : solid state　　(l) : liquid state (g) : gas state　　(aq) : aqueous solution	$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$　　(Balanced)

Some useful hints in balancing chemical equations:

1. Never change the chemical formulae of the reactants or products in order to get the equation balanced. You can only add correct coefficients in order to an equation balanced.

2. You can start balancing the equation from any one chemical in the equation. It may be a reactant or a product. Usually, it will go easy if you start from the most complicated chemical in the equation.

3. The first coefficient is invented by you. If you have no idea, you can start from '1'. Don't be afraid if you end up with some fractional coefficients during the course of balancing. They can be converted to a whole number at the end if you multiply all the coefficients by a suitable whole number. Our practice is to express the coefficients in the simplest integral ratio.

4. How to count the atoms in case of compounds consisting of polyatomic ions? It can be explained by the example equation below:

$$3H_2SO_4 + 2Fe(OH)_3 \longrightarrow Fe_2(SO_4)_3 + 6H_2O$$
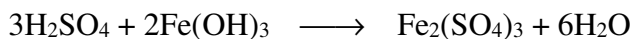
Number of H atoms from H_2SO_4 in the reactants = 3 x 2 = 6

Number of H atoms from $Fe(OH)_3$ in the reactants = 2 x 3 = 6

∴ Total number of H atoms in the reactants = 6 + 6 = 12

Can you count the total number of O atoms in the reactants?

More examples on writing balanced chemical equations

Example Write a balanced chemical equation for the reaction between calcium and dilute hydrochloric acid.

Solution:

Calcium and dilute hydrochloric acid react to give calcium chloride solution and hydrogen gas.

1 Write down the word equation.	calcium + hydrochloric acid $\longrightarrow$ calcium chloride + hydrogen
2 Write down the chemical formulae of the reactants and products.	$Ca + HCl \longrightarrow CaCl_2 + H_2$
3 To balance the chlorine atom, put the coefficient '2' before HCl. Now the chemical equation is balanced.	$Ca + 2HCl \longrightarrow CaCl_2 + H_2$ (Balanced)
4 Write down the state symbol after each substance.	$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$ (Balanced)

Example

Write a balanced chemical equation for the reaction between silver nitrate solution and zinc. (Hint: silver nitrate solution will react with zinc to form silver and zinc nitrate solution.)

Solution:

Silver nitrate solution reacts with zinc to give silver and zinc nitrate solution.

1 Write down the word equation.	silver nitrate + zinc $\longrightarrow$ silver + zinc nitrate
2 Write down the chemical formulae of the reactants and products.	$AgNO_3 + Zn \longrightarrow Ag + Zn(NO_3)_2$
3 a) To balance the polyatomic ion NO_3^-, put the coefficient '2' before $AgNO_3$.	$2AgNO_3 + Zn \longrightarrow Ag + Zn(NO_3)_2$
b) To balance the silver atom, put the coefficient '2' before Ag.	$2AgNO_3 + Zn \longrightarrow 2Ag + Zn(NO_3)_2$ (Balanced)
4 Write down the state symbol after each substance.	$2AgNO_3(aq) + Zn(s) \longrightarrow 2Ag(s) + Zn(NO_3)_2(aq)$ (Balanced)

For more exercises, refer to the exercises on balancing chemical equations in the worksheets.

What determines the reactivity of a metal?

Metals actually lose electrons to form cations in their chemical reactions. The reactivity of a metal is actually determined by the ease of formation of the cation. The easier the cation is formed, the more **reactive** the metal will be.

For metals within the same group, it is found that reactivity always increase down the group. It is because the metal atoms are becoming larger and larger as we move down the group. The outermost shell electrons will be held less strongly by the nuclei if the atoms are becoming large as they will be further and further away from the nuclei. This explains the following order of reactivity: reactivity of K > reactivity of Na   and reactivity of Ca > reactivity of Mg

For metals within the same period, it is found that the metal in Group I is more reactive than the one in Group II and is in turn more reactive than the one in Group III.

Group I metals will form ions of charge +1. Only 1 electron is required to remove in their reactions. Group II metals will form ions of charge +2. Two electrons are required to remove in their reactions. Group III metals will form ions of charge +3. Three electrons are required to remove in their reactions. The larger the number of electrons has to be removed, the more

difficult it will be and the less reactive would be the metal.

Trends in reactivity of metals from the periodic table

readiness to lose electrons increasing

reactivity increasing

readiness to lose electrons increasing

reactivity increasing

Group		
I	II	III

Period 2 — Li, Be

Period 3 — Na, Mg, Al

Period 4 — K, Ca
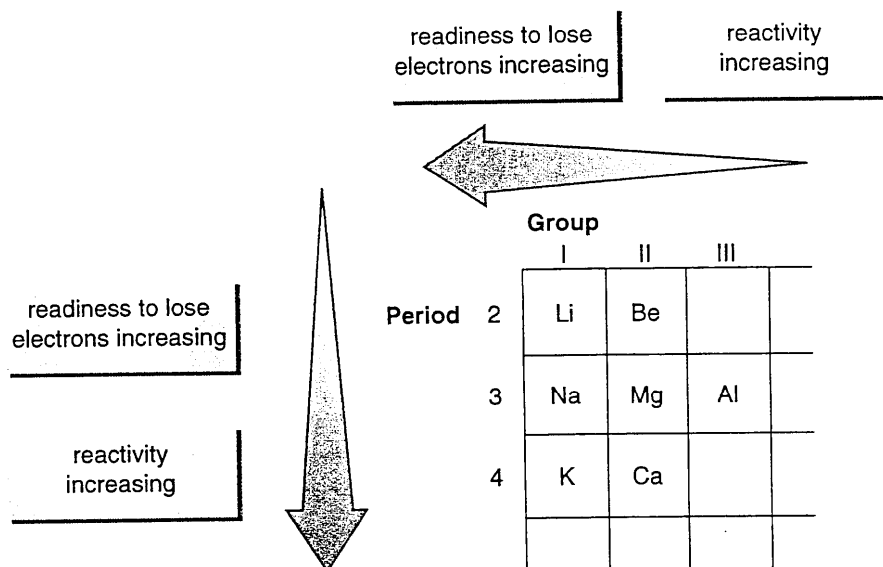
Displacement reactions

When we dip a piece of zinc into a blue copper(II) sulphate solution, the zinc metal slowly becomes coated with a thick brown layer of copper. The blue colour of the solution fades gradually. However, when we place a piece of copper into a zinc sulphate solution, no reaction takes place.

A reaction in which one element displaces another element from a compound is called a **displacement reaction**. In the above reaction, zinc displaces copper from copper(II) sulphate solution.

> **A metal will displace a less reactive metal from a solution of the compound of the less reactive metal, but not vice versa.**

In the reaction between zinc and copper(II) sulphate solution, zinc displaces copper from copper(II) sulphate solution to form zinc sulphate solution and copper deposits.

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$
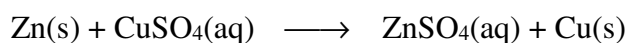
Zinc is more reactive than copper. The above reaction will occur. However, if copper is placed into zinc sulphate solution, no reaction will be seen. Copper cannot displace zinc from zinc sulphate solution as it is less reactive than zinc.

Similarly, as copper is more reactive than silver, it can displace silver from silver nitrate solution. When a piece of copper is put into silver nitrate solution, a silvery deposit is gradually seen and the solution gradually turns blue.

$$Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$
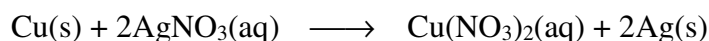
Ionic equations

Consider the reaction between zinc and copper(II) sulphate solution. We can represent the reaction by the following equation:

$$Zn(s) + CuSO_4(aq) \longrightarrow ZnSO_4(aq) + Cu(s)$$
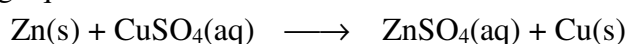
In this reaction, each zinc atom loses 2 electrons to form a zinc ion. Its change can be represented as: $Zn(s) \longrightarrow Zn^{2+}(aq) + 2e^-$

Each copper(II) ion gains 2 electrons to form a copper atom. Its change can be represented as:
$Cu^{2+}(aq) + 2e^- \longrightarrow Cu(s)$

To sum up the changes, the equations can be simplified as one called ionic equation:

$$Zn(s) + Cu^{2+} \longrightarrow Zn^{2+}(aq) + Cu(s)$$
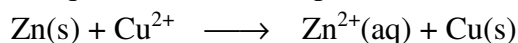
An equation like this, which involves only ions formed or changed during a reaction, is called an **ionic equation**. The main advantage of an ionic equation is that it shows clearly which ions are taking part in a reaction.

An ionic equation is a special type of chemical equation. Just like common chemical equation, it is also balanced. Actually, apart from the fact that the total number of each type of atoms on the reactants side should be equal to that on the products side, the total charges on the reactants side should be equal to those of the products side as well.

Steps for writing ionic equations

To write an ionic equation of a reaction, you could start from the chemical equation and follow the following steps. Let us take the reaction between copper and silver nitrate solution as an example.

Step	Example
1 Write down the word equation.	copper + silver nitrate $\longrightarrow$ copper(II) nitrate + silver
2 Write down the chemical formulae of the reactants and products.	$Cu + AgNO_3 \longrightarrow Cu(NO_3)_2 + Ag$
3 Balance the chemical equation and write down the state symbol after each substance.	$Cu(s) + 2AgNO_3(aq) \longrightarrow Cu(NO_3)_2(aq) + 2Ag(s)$ **(Balanced)**
4 Decide which substances exist as mobile ions in solution. Write down the ions of these substances. (Do not change the chemical formulae of solids and gases.)	Silver nitrate and copper(II) nitrate are both ionic compounds. They exist as mobile ions in solution. Silver nitrate solution contains $Ag^+(aq)$ and $NO_3^-(aq)$. Copper(II) nitrate solution contains $Cu^{2+}(aq)$ and $NO_3^-(aq)$. We can rewrite the above equation as follows: $Cu(s) + 2Ag^+(aq) + \cancel{2NO_3^-(aq)}$ $\longrightarrow Cu^{2+}(aq) + \cancel{2NO_3^-(aq)} + 2Ag(s)$ **(Balanced)**
5 Delete ions that are not taking part in the reaction to obtain the ionic equation.	$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$ **(Balanced)** (net charge = +2) (net charge = +2)

Exercises

1.   Write an ionic equation for each of the following reactions:
     (a) Zinc is added to copper(II) chloride solution

     (b) Magnesium is added to silver nitrate solution.

2. In an experiment, an iron nail was placed in a beaker of copper(II) sulphate solution.

    iron nail

    copper(II) sulphate solution

    (a) State the expected observations

    (b) Write a chemical equation for the reaction involved.

    (c) Write an ionic equation for the reaction involved.

3. Consider the following cases:

    lead(II) nitrate
    solution

    zinc

    A

    iron(II) nitrate
    solution

    magnesium

    B

    For each case,
    (a) predict whether a reaction will take place ; and
         Case A: Case B:

    (b) write a chemical equation and an ionic equation if a reaction takes place.
         Chemical equation:

         Ionic equation:

Relationship between the extraction method and position of metals in the reactivity series

The table below summarizes the extraction methods of some metals in the reactivity series. Notice the following points:

1. metals at top of the reactivity series (potassium, sodium, magnesium and aluminium) are extracted by electrolysis;
2. metals in the middle (zinc, iron, lead and also copper if extracted from its oxides) are extracted by reduction of their oxides with carbon;
3. metals near the bottom (copper and copper if extracted from its sulphide) are extracted by heating in the air; and
4. metals at the bottom (silver and gold) are extracted by physical methods.

(The table below is for your reference only)

|	Metal	Ore	Main metallic compound in the ore	Extraction method
most difficult	Potassium (K)	*carnallite*	hydrated potassium magnesium chloride (KCl•MgCl$_2$•6H$_2$O)	
	Sodium (Na)	*rock salt*	sodium chloride (NaCl)	electrolysis of molten ore
	Magnesium (Mg)	*magnesite*	magnesium carbonate (MgCO$_3$)	
	Aluminium (Al)	*bauxite*	hydrated aluminium oxide (Al$_2$O$_3$•2H$_2$O)	
	Zinc (Zn)	*zinc blende*	zinc sulphide (ZnS)	sulphide $\xrightarrow[\text{in air}]{\text{heat}}$ oxide then oxide $\xrightarrow[\text{with carbon}]{\text{heat}}$ metal
ease of extraction	Iron (Fe)	*haematite*	iron(III) oxide (Fe$_2$O$_3$)	oxide $\xrightarrow[\text{with carbon}]{\text{heat}}$ metal
	Lead (Pb)	*galena*	lead(II) sulphide (PbS)	similar to that of Zn
	Copper (Cu)	*copper pyrite*	copper(II) iron(II) sulphide (CuFeS$_2$)	sulphide $\xrightarrow[\text{in air}]{\text{controlled heat}}$ metal
	Mercury (Hg)	cinnabar	mercury(II) sulphide (HgS)	sulphide $\xrightarrow[\text{in air}]{\text{heat}}$ metal
	Silver (Ag)	*argentite* or as free element	silver sulphide (Ag$_2$S) or free element of silver	displacement from solution or mechanical separation
easiest	Gold (Au)	as free element	free element of gold	mechanical separation

Prediction of metal reactions using the reactivity series
Reactivity series and reduction of metal oxides
Reduction means the removal of oxygen atoms from a chemical containing oxygen atoms(e.g. an oxide) in order to release the other element.

For instance, metal low in the metal reactivity series does not burn in air to form oxide. It means that the oxide of the metal will be unstable and can be reduced easily, even by simple heating.
e.g. Silver oxide can be reduced to silver metal by simple heating.
$$2Ag_2O(s) \longrightarrow 4Ag(s) + O_2(g)$$
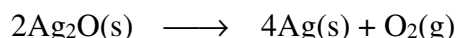
Metals in the middle of the reactivity series are more difficult to be reduced. Their oxides have to be reduced by heating with charcoal (a form of carbon).
e.g. Lead(II) oxide can be reduced by the set up below.
$$2PbO(s) + C(s) \longrightarrow 2Pb(s) + CO_2(g)$$
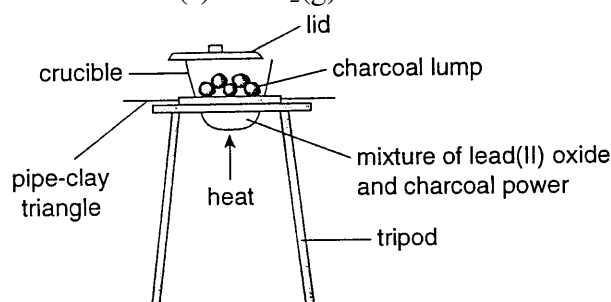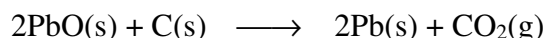
In general, the trend in the ease of reduction of metal oxides is exactly the opposite to the trend in the reactivity of metal. The more reactive the metal is, the more difficult its oxide is to be reduced. Refer to the diagram below.
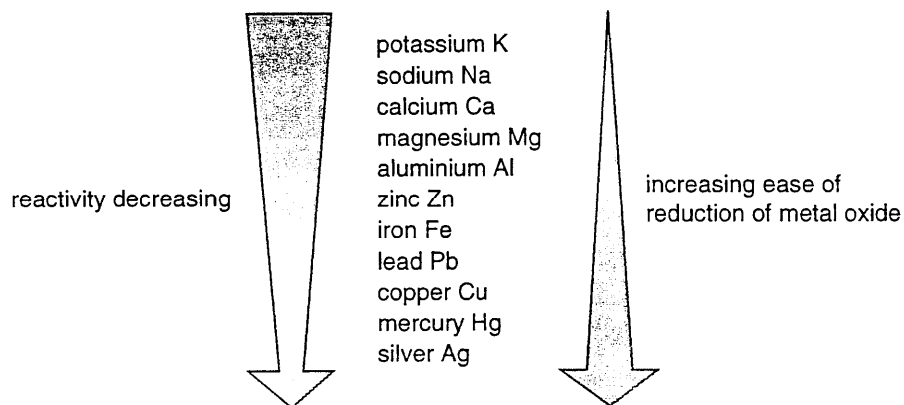
Distinguishing metals by their reactions
Example
X, Y and Z are three different metals. The table below shows the results of two experiments carried out using the metals.

Experiment	Metal		
	X	Y	Z
Adding he metal to cold water	A colourless gas was given off	No observable change	No observable change
Adding the metal to copper(II) sulphate solution	A colourless gas was given off and a reddish brown deposit formed	No observable change	A reddish brown deposit formed

(a) Name the colourless gas given off when X was added to cold water.

(b) Name the type of reaction that occurred when Z was added to copper(II) sulphate solution.
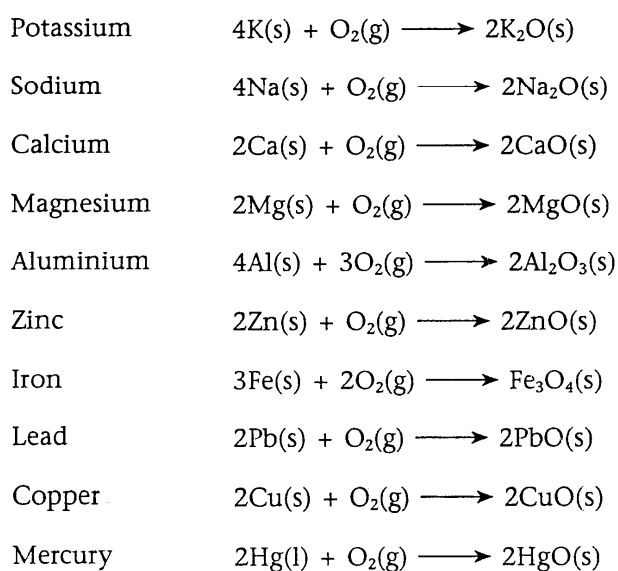
(c) Arrange the three metals X, Y and Z in order of increasing reactivity. Explain your answer.

(d) Predict whether X or Y was discovered earlier. Explain your answer.

(e) Explain why a colourless gas was given of when X was added to copper(II) sulphate solution.

Solution

(a) Hydrogen

(b) Displacement reaction

(c) Reactivity: $Y < Z < X$

X was the **most reactive** metal because **only** X could react with cold water.

Y was the **least reactive** metal because **only** Y had could not displace copper from copper(II) sulphate solution.

(d) Y was discovered earlier. It is less reactive than X and thus easier to extract.

(e) X is a very reactive metal. It reacted with the water in the copper(II) sulphate solution to give hydrogen gas.

Exercise

The table below lists some information about three metals, X, Y and Z.

Metal	X	Y	Z
Atomic number	12	-	-
Reaction with dilute hydrochloric acid	No information	No observable change	No observable change
Heating the metal oxide	No observable change	A solid with metallic luster formed	No observable change

(a) Name metal X.

(b) X was added to dilute hydrochloric acid.
   (i) What will be observed?

   (ii) Write a chemical equation for the reaction involved.

(c) Draw electron diagrams for the TWO products formed in (b) above, showing electrons in the outermost shell only.

(d) Arrange the three metals in order of increasing reactivity. Explain your answer.

Unit Summary

1 Metals can be arranged in a list of decreasing reactivity. This order is called the reactivity series.

2 The following table summarizes the reactions of metals in the reactivity series.

Metal	Reaction of metal with air or oxygen	Reaction of metal with water or steam	Reaction of metal with dilute acids (dilute HCl and H₂SO₄)	Displacement reaction
Potassium	burn vigorously	react with cold water	react explosively	K, Na and Ca react with water to give off hydrogen gas
Sodium		metal + water		
Calcium	react with decreasing vigour	⟶ metal hydroxide + hydrogen	metal + dilute hydrochloric acid	
Magnesium		react with steam	⟶ metal chloride + hydrogen	a metal will always displace a less reactive metal from a solution of the compound of the less reactive metal, but not vice versa
Aluminium		metal + steam		
Zinc		⟶ metal oxide + hydrogen	metal + dilute sulphuric acid	
Iron			⟶ metal sulphate + hydrogen	
Lead	a layer of oxide forms on the surface			
Copper				
Mercury				
Silver		no reaction	no reaction	
Platinum	no reaction			
Gold				

3 Balanced chemical equations for reactions of some common metals with oxygen in the air:

Potassium $4K(s) + O_2(g) \longrightarrow 2K_2O(s)$

Sodium $4Na(s) + O_2(g) \longrightarrow 2Na_2O(s)$

Calcium $2Ca(s) + O_2(g) \longrightarrow 2CaO(s)$

Magnesium $2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$

Aluminium $4Al(s) + 3O_2(g) \longrightarrow 2Al_2O_3(s)$

Zinc $2Zn(s) + O_2(g) \longrightarrow 2ZnO(s)$

Iron $3Fe(s) + 2O_2(g) \longrightarrow Fe_3O_4(s)$

Lead $2Pb(s) + O_2(g) \longrightarrow 2PbO(s)$

Copper $2Cu(s) + O_2(g) \longrightarrow 2CuO(s)$

Mercury $2Hg(l) + O_2(g) \longrightarrow 2HgO(s)$

5   Balanced equations for reactions of some common metals with dilute hydrochloric acid and dilute sulphuric acid:

Metal	Chemical equation	Ionic equation
Calcium	$Ca(s) + 2HCl(aq) \longrightarrow CaCl_2(aq) + H_2(g)$	$Ca(s) + 2H^+(aq) \longrightarrow Ca^{2+}(aq) + H_2(g)$
	$Ca(s) + H_2SO_4(aq) \longrightarrow CaSO_4(s) + H_2(g)$	—
Magnesium	$Mg(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + H_2(g)$	$Mg(s) + 2H^+(aq) \longrightarrow Mg^{2+}(aq) + H_2(g)$
	$Mg(s) + H_2SO_4(aq) \longrightarrow MgSO_4(aq) + H_2(g)$	
Aluminium	$2Al(s) + 6HCl(aq) \longrightarrow 2AlCl_3(aq) + 3H_2(g)$	$2Al(s) + 6H^+(aq) \longrightarrow 2Al^{3+}(aq) + 3H_2(g)$
	$2Al(s) + 3H_2SO_4(aq) \longrightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$	
Zinc	$Zn(s) + 2HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$	$Zn(s) + 2H^+(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$
	$Zn(s) + H_2SO_4(aq) \longrightarrow ZnSO_4(aq) + H_2(g)$	
Iron	$Fe(s) + 2HCl(aq) \longrightarrow FeCl_2(aq) + H_2(g)$	$Fe(s) + 2H^+(aq) \longrightarrow Fe^{2+}(aq) + H_2(g)$
	$Fe(s) + H_2SO_4(aq) \longrightarrow FeSO_4(aq) + H_2(g)$	

6   Useful information obtained from a chemical equation include

a)   the reactant(s) involved;

b)   the product(s) formed;

c)   the states of substances involved; and

d)   the number of particles of each type of substance involved.

7   Atoms of more reactive metals lose outermost shell electrons to form cations more readily.

8   Ionic equations for some displacement reactions:

Metal	Reaction with solution of	Ionic equation
Magnesium	aluminium compound e.g. $AlCl_3$, $Al_2(SO_4)_3$, $Al(NO_3)_3$	$3Mg(s) + 2Al^{3+}(aq) \longrightarrow 3Mg^{2+}(aq) + 2Al(s)$
	zinc compound e.g. $ZnCl_2$, $ZnSO_4$, $Zn(NO_3)_2$	$Mg(s) + Zn^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Zn(s)$
	iron(II) compound e.g. $FeCl_2$, $FeSO_4$, $Fe(NO_3)_2$	$Mg(s) + Fe^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Fe(s)$
	copper(II) compound e.g. $CuCl_2$, $CuSO_4$, $Cu(NO_3)_2$	$Mg(s) + Cu^{2+}(aq) \longrightarrow Mg^{2+}(aq) + Cu(s)$
	silver compound e.g. $AgNO_3$	$Mg(s) + 2Ag^+(aq) \longrightarrow Mg^{2+}(aq) + 2Ag(s)$
Zinc	iron(II) compound	$Zn(s) + Fe^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Fe(s)$
	copper(II) compound	$Zn(s) + Cu^{2+}(aq) \longrightarrow Zn^{2+}(aq) + Cu(s)$
	silver compound	$Zn(s) + 2Ag^+(aq) \longrightarrow Zn^{2+}(aq) + 2Ag(s)$
Iron	copper(II) compound	$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$
	silver compound	$Fe(s) + 2Ag^+(aq) \longrightarrow Fe^{2+}(aq) + 2Ag(s)$
Copper	silver compound	$Cu(s) + 2Ag^+(aq) \longrightarrow Cu^{2+}(aq) + 2Ag(s)$

9   The lower the position of the metal in the reactivity series, the more easily it can be extracted from its ores.