9 The Periodic Table

FOCUS POINTS

★ How can the Periodic Table help us to understand the way chemical elements react?
★ What are the trends in properties and reactivity for different groups in the Periodic Table?

In this chapter, we will look at the way in which all the chemical elements have been organised into a table that can help to predict their properties and their chemical reactions. Chemists began to try to organise the elements around 150 years ago but found it very difficult and, as new elements were discovered, they had to revise their attempts. Eventually the format of the Periodic Table which we see today was produced and, as new elements were discovered, they fitted in perfectly.

The aim of this chapter is to help you to use the Periodic Table to understand chemistry more easily. For example, if you know some of the properties of sodium, the Periodic Table can help you to predict the properties of francium.

By the end of this chapter you will have learned the trends in properties of four groups of elements and you should understand how useful the Periodic Table is to your study of chemistry.

9.1 Development of the Periodic Table

The Periodic Table is a vital tool used by chemists to predict the way in which elements react during chemical reactions. It is a method of categorising elements according to their properties. Scientists started to look for a way in which to categorise the known elements around 150 years ago.

The Periodic Table was devised in 1869 by the Russian Dmitri Mendeleev, who was the Professor of Chemistry at St Petersburg University (Figure 9.1). His Periodic Table was based on the chemical and physical properties of the 63 elements that had been discovered at that time.

Other scientists had attempted to categorise the known elements from the early 19th century but Mendeleev's classification proved to be the most successful.

Going further

Mendeleev arranged all the 63 known elements in order of increasing atomic weight but in such a way that elements with similar properties were in the same vertical column. He called the vertical columns **groups** and the horizontal rows **periods** (Figure 9.2). If necessary, he left gaps in the table.

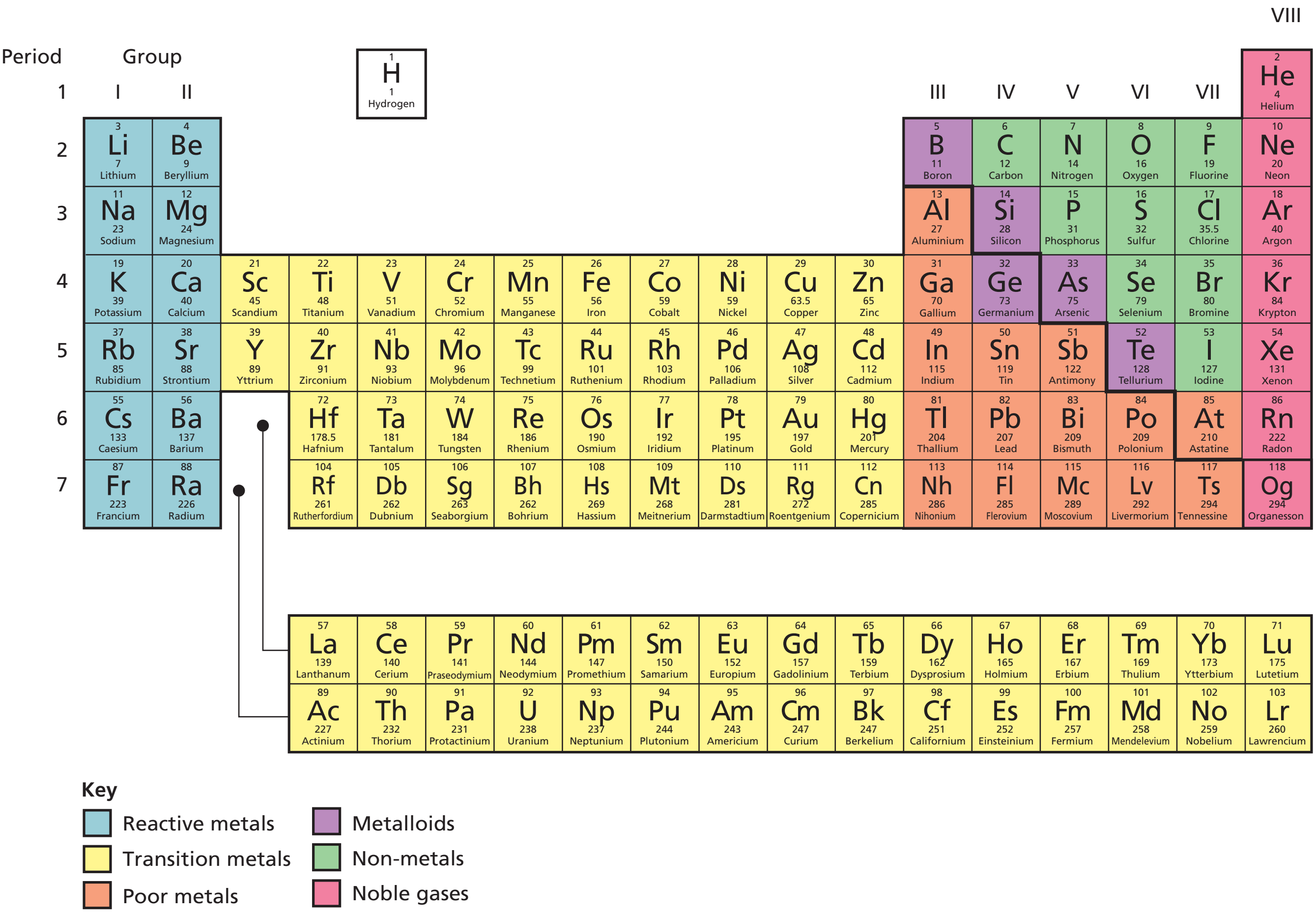
As a scientific idea, Mendeleev's Periodic Table was tested by making predictions about elements that were unknown at that time but could possibly fill the gaps. Three of these gaps are shown by the symbols * and † in Figure 9.2. As new elements were discovered, they were found to fit easily into the classification. For example, Mendeleev predicted the properties of the missing element 'eka-silicon' (†). He predicted the colour, density and melting point as well as its atomic weight.

In 1886 the element we now know as germanium was discovered in Germany by Clemens Winkler; its properties were almost exactly those Mendeleev had predicted. In all, Mendeleev predicted the atomic weight of ten new elements, of which seven were eventually discovered – the other three, atomic weights 45, 146 and 175, do not exist!

▲ **Figure 9.1** Dmitri Mendeleev (1834–1907)

Period	Group							
	1	2	3	4	5	6	7	8
1	H							
2	Li	Be	B	C	N	O	F	
3	Na	Mg	Al	Si	P	S	Cl	
4	K	Ca	*	Ti	V	Cr	Mn	Fe Co Ni
	Cu	Zn	*	†	As	Se	Br	

▲ **Figure 9.2** Mendeleev's Periodic Table. He left gaps for undiscovered elements



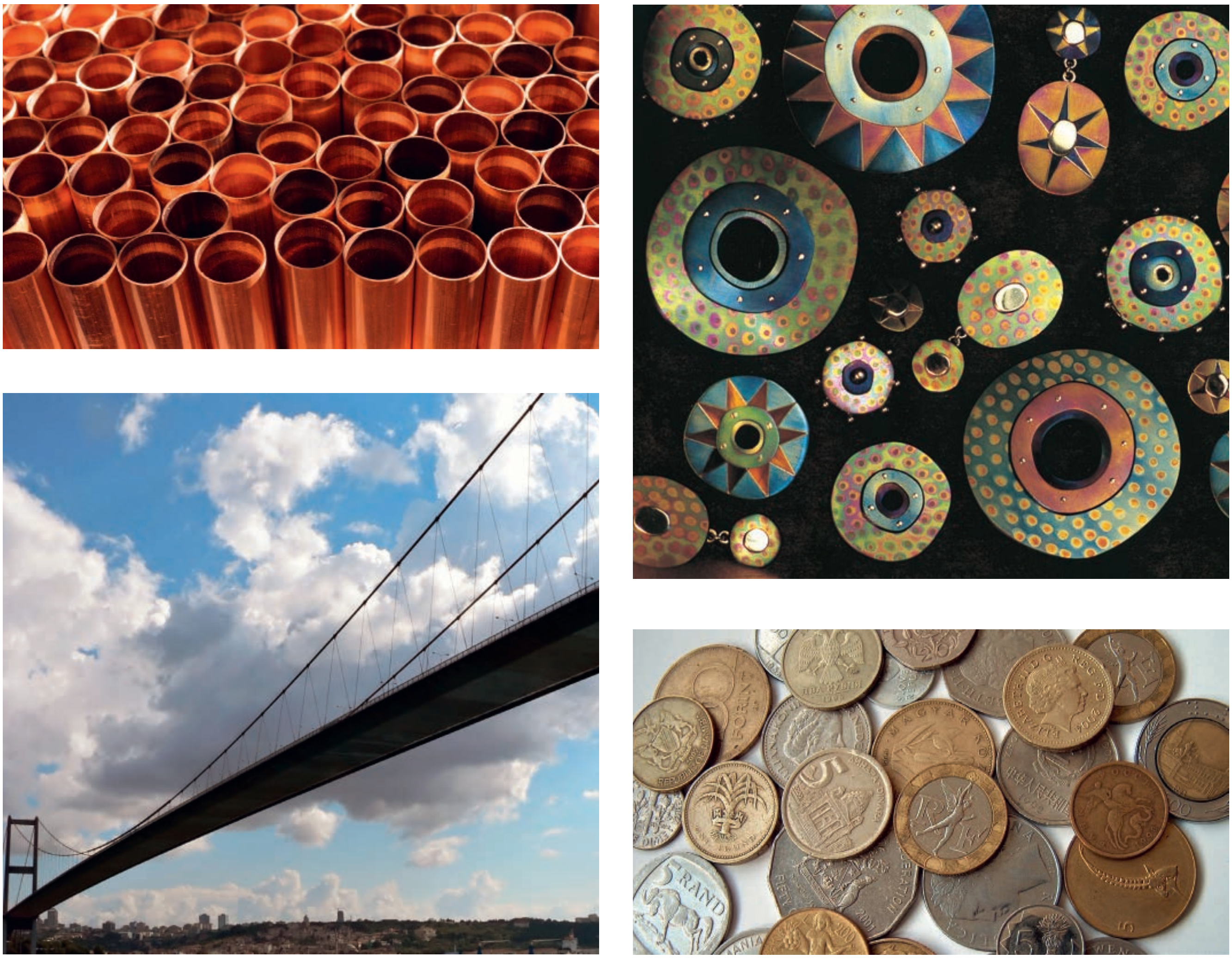
▲ **Figure 9.3** The modern Periodic Table

The success of Mendeleev's predictions showed that his ideas were probably correct. His Periodic Table was quickly accepted by scientists as an important summary of the properties of the elements.

Mendeleev's Periodic Table has been modified in the light of work carried out by Ernest Rutherford and Henry Moseley. Discoveries about sub-atomic particles led them to realise that the elements should be arranged by proton number. In the modern Periodic Table, the 118 known elements are arranged in order of increasing proton number (Figure 9.3).

Those elements with similar chemical properties are found in the same columns or **groups**. There are eight groups of elements. The first column is called Group I, the second Group II, and so on up to Group VII. The final column in the Periodic Table is called Group 0 (or Group VIII). Some of the groups have been given names.

Group I: The **alkali metals**
Group II: The **alkaline earth metals**
Group VII: The **halogens**
Group 0: Inert gases or **noble gases**



▲ **Figure 9.4** Transition elements have a wide range of uses, both as elements and as alloys

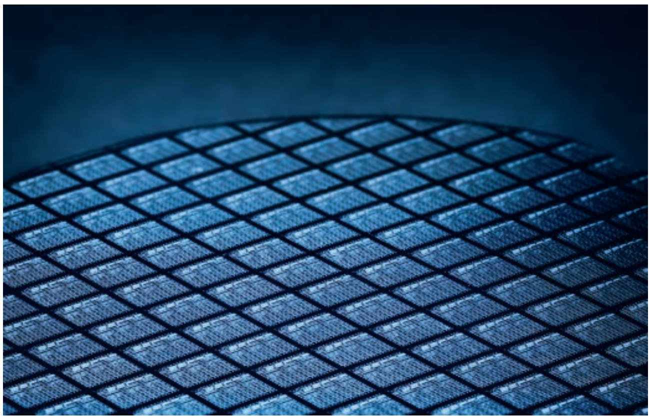
The horizontal rows are called **periods** and these are numbered 1–7 going down the Periodic Table.

Between Groups II and III is the block of elements known as the **transition elements** (Figure 9.4).

The Periodic Table can be divided into two as shown by the bold line that starts beneath boron, in Figure 9.3. The elements to the left of this line are metals (fewer than three-quarters) and those on the right are non-metals (fewer than one-quarter). The elements which lie on this dividing line are known as **metalloids** (Figure 9.5). The metalloids behave in some ways as metals and in others as non-metals.

If you look at the properties of the elements across a period of the Periodic Table, you will notice certain trends. For example, there is:

- a gradual change from metal to non-metal
- an increase in the number of electrons in the outer electron shell of the element
- a change in the structure of the element, from giant metallic in the case of metals (e.g. magnesium, p. 50, Figure 3.34), through giant covalent (e.g. diamond, p. 46, Figure 3.29), to simple molecular (e.g. chlorine, p. 39, Figure 3.13).



▲ **Figure 9.5** The metalloid silicon is used to make silicon 'chips'

9.2 Electronic configuration and the Periodic Table

The number of electrons in the outer electron shell is discussed in Chapter 2 (p. 26). This corresponds with the number of the group in the Periodic Table in which the element is found. For example, the elements shown in Table 9.1 have one electron in their outer electron shell and they are all found in Group I. The elements in Group 0, however, are an exception to this rule, as they have two or eight electrons in their outer electron shell. The outer electrons are mainly responsible for the chemical properties of any element and, therefore, elements in the same group have similar chemical properties (Tables 9.2 and 9.3). The number of occupied shells is equal to the period number in which that element is found. For example, we know that sodium is in Group I because it has only one electron in its outer shell but we can also say that it is in Period 3 of the Periodic Table as it has electrons in the first three shells.

▼ **Table 9.1** Electronic configuration of the first three elements of Group I

Element	Symbol	Proton number	Electronic configuration
Lithium	Li	3	2,1
Sodium	Na	11	2,8,1
Potassium	K	19	2,8,8,1

▼ **Table 9.2** Electronic configuration of the first three elements of Group II

Element	Symbol	Proton number	Electronic configuration
Beryllium	Be	4	2,2
Magnesium	Mg	12	2,8,2
Calcium	Ca	20	2,8,8,2

▼ **Table 9.3** Electronic configuration of the first three elements in Group VII

Element	Symbol	Proton number	Electronic configuration
Fluorine	F	9	2,7
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7

The metallic character of the elements in a group increases as you move down the group. This is because electrons become easier to lose as the outer shell electrons become further from the nucleus. There is less attraction between the nucleus and the outer shell electrons because of the increased distance between them.

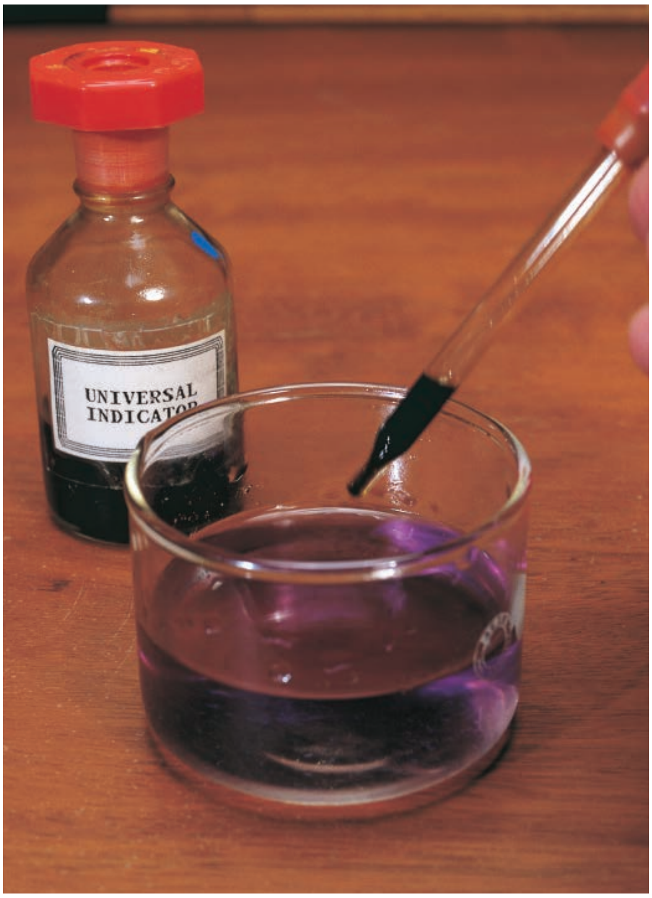
9.3 Group I – the alkali metals

Group I consists of the five metals lithium, sodium, potassium, rubidium and caesium, and the radioactive element francium. Lithium, sodium and potassium are commonly available for use in school. They are all very reactive metals and they are stored under oil to prevent them coming into contact with water or air. These three metals have the following properties.

- They are good conductors of electricity and heat.
- They are soft metals. Lithium is the hardest and potassium the softest.
- They are metals with low densities. For example, lithium has a density of 0.53 g/cm^3 and potassium has a density of 0.86 g/cm^3.
- They have shiny surfaces when freshly cut with a knife (Figure 9.6).

▲ **Figure 9.6** Cutting sodium metal

a Potassium reacts very vigorously with cold water



b An alkaline solution is produced when potassium reacts with water

▲ **Figure 9.7**

- They have low melting points. For example, lithium has a melting point of 181°C and potassium has a melting point of 64°C.
- They burn in oxygen or air, with characteristic flame colours, to form white solid oxides. For example, lithium reacts with oxygen in air to form white lithium oxide, according to the following equation:

lithium + oxygen → lithium oxide

$$4Li(s) + O_2(g) \rightarrow 2Li_2O(s)$$

These Group I oxides all react with water to form alkaline solutions of the metal hydroxide.

lithium oxide + water → lithium hydroxide

$Li_2O(s) + H_2O(l) \rightarrow 2LiOH(aq)$

- They react vigorously with water to give an alkaline solution of the metal hydroxide as well as producing hydrogen gas. For example:

potassium + water → potassium hydroxide + hydrogen gas

$2K(s) + 2H_2O(l) \rightarrow 2KOH(aq) + H_2(g)$

- Of the first three metals in Group I, potassium is the most reactive towards water (Figure 9.7), followed by sodium and then lithium. Such gradual changes we call **trends**. Trends are useful to chemists as they allow predictions to be made about elements we have not observed in action.
- They react vigorously with halogens, such as chlorine, to form metal halides, for example, sodium chloride (Figure 9.8).

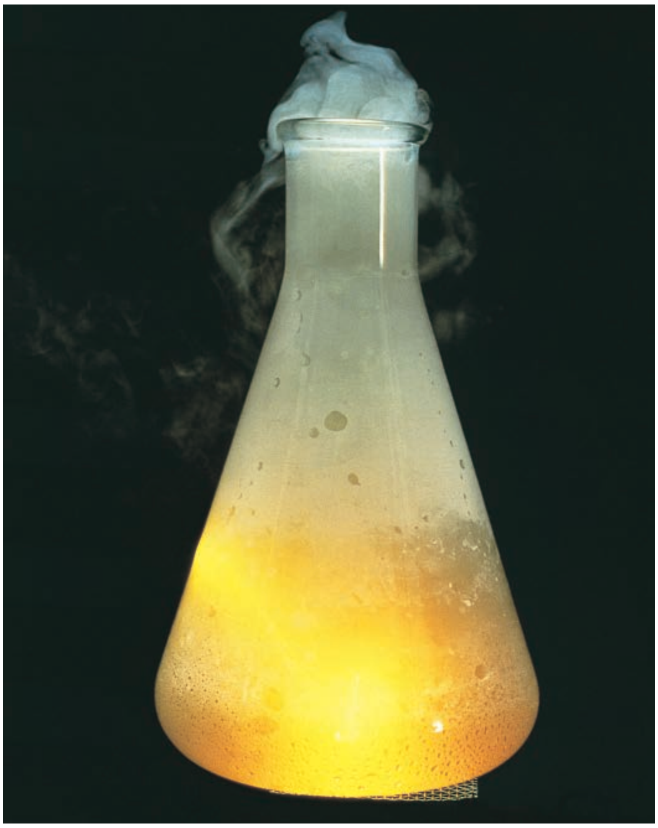
sodium + chlorine → sodium chloride

$2Na(s) + Cl_2(g) \rightarrow 2NaCl(s)$

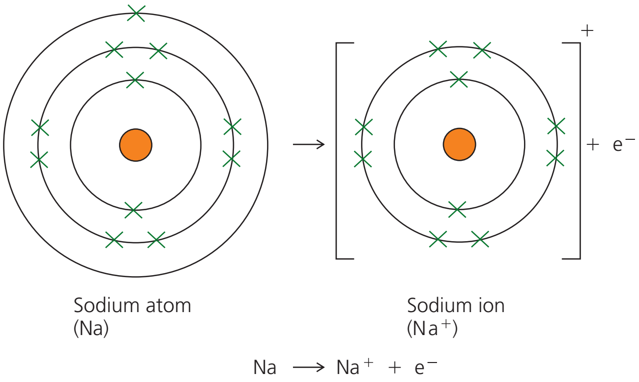
Considering the group as a whole, the further down the group you go, the more reactive the metals become. Francium is, therefore, the most reactive Group I metal.

Table 9.1 shows the electronic configuration of the first three elements of Group I. You will notice in each case that the outer electron shell contains only one electron. When these elements react, they lose this outer electron and, in doing so, become more stable because they obtain the electronic configuration of a noble gas. You will learn more about the stable nature of these gases later in this chapter.

When, for example, the element sodium reacts, it loses its outer electron. This requires energy to overcome the electrostatic attractive forces between the outer electron and the positive nucleus (Figure 9.9).

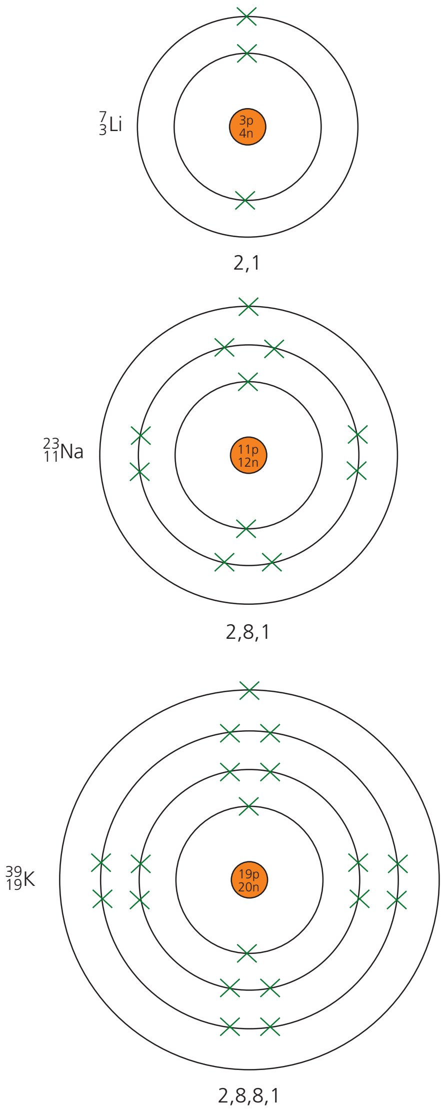


▲ **Figure 9.8** A very vigorous reaction takes place when sodium burns in chlorine gas. Sodium chloride is produced



▲ **Figure 9.9** This sodium atom loses an electron to become a sodium ion

Look at Figure 9.10. Why do you think potassium is more reactive than lithium or sodium?



▲ **Figure 9.10** Electronic configuration of lithium, sodium and potassium

Potassium is more reactive because less energy is required to remove the outer electron from its atom than for lithium or sodium. This is because as you go down the group, the size of the atoms increases and the outer electron gets further away from the nucleus and becomes easier to remove.

Test yourself

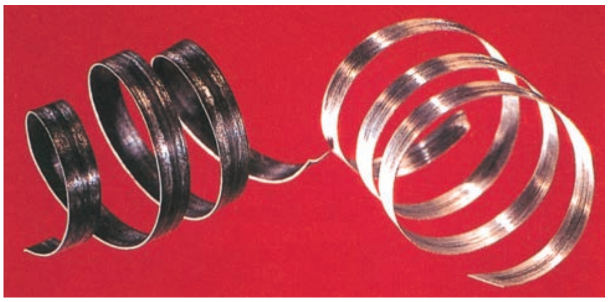
1 Write word and balanced chemical equations for the reactions between:
   a sodium and oxygen
   b sodium and water.
2 a Using the information on p. 139, predict the properties of the element francium related to its melting point, density and softness.
   b Predict how francium would react with water and write a balanced equation for the reaction.
3 Write word and balanced chemical equations for the reactions between:
   a magnesium and water
   b calcium and oxygen.
4 Explain the fact that calcium is more reactive than magnesium in terms of their electronic configurations.

Going further

Group II – the alkaline earth metals

Group II consists of the five metals beryllium, magnesium, calcium, strontium and barium, and the radioactive element radium. Magnesium and calcium are generally available for use in school. These metals have the following properties.

- They are harder than those in Group I.
- They are silvery-grey in colour when pure and clean. They tarnish quickly, however, when left in air due to the formation of a metal oxide on their surfaces (Figure 9.11).



▲ **Figure 9.11** Tarnished (left) and cleaned-up magnesium

- They are good conductors of heat and electricity.
- They burn in oxygen or air with characteristic flame colours to form solid white oxides. For example:

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

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

- They react with water, but they do so much less vigorously than the elements in Group I. For example:

$$\text{calcium} + \text{water} \rightarrow \text{calcium hydroxide} + \text{hydrogen gas}$$

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(aq) + H_2(g)$$

Considering the group as a whole, the further down the group you go, the more reactive the elements become.

9.4 Group VII – the halogens

Group VII consists of the four elements fluorine, chlorine, bromine and iodine, and the radioactive element astatine. Of these five elements, chlorine, bromine and iodine are generally available for use in school.

- These elements are coloured and become darker going down the group (Table 9.4).
- They exist as diatomic molecules, for example, Cl_2, Br_2 and I_2.
- At room temperature and pressure they show a gradual change from a gas (Cl_2), through a liquid (Br_2), to a solid (I_2) (Figure 9.12) as the density increases.
- They form molecular compounds with other non-metallic elements, for example HCl.
- They react with hydrogen to produce the hydrogen halides, which dissolve in water to form acidic solutions.

$$\text{hydrogen} + \text{chlorine} \rightarrow \text{hydrogen chloride}$$

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

$$\text{hydrogen chloride} + \text{water} \rightarrow \text{hydrochloric acid}$$

$$HCl(g) + H_2O \rightarrow HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$$

- They react with metals to produce ionic metal halides, for example, chlorine and iron produce iron(III) chloride.

$$\text{iron} + \text{chlorine} \rightarrow \text{iron(III) chloride}$$

$$2Fe(s) + 3Cl_2(g) \rightarrow 2FeCl_3(s)$$

a Chlorine, bromine and iodine

b Chlorine gas bleaches moist indicator paper

▲ **Figure 9.12**

▼ **Table 9.4** Colours of some halogens

Halogen	Colour
Chlorine	Pale yellow-green gas
Bromine	Red-brown liquid
Iodine	Grey-black solid

Displacement reactions

If chlorine is bubbled into a solution of potassium iodide, the less reactive halogen, iodine, is displaced by the more reactive halogen, chlorine, as you can see from Figure 9.13:

potassium iodide + chlorine → potassium chloride + iodine

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

▲ **Figure 9.13** Iodine being displaced from potassium iodide solution as chlorine is bubbled through

The observed order of reactivity of the halogens, confirmed by similar **displacement reactions**, is:

decreasing reactivity →

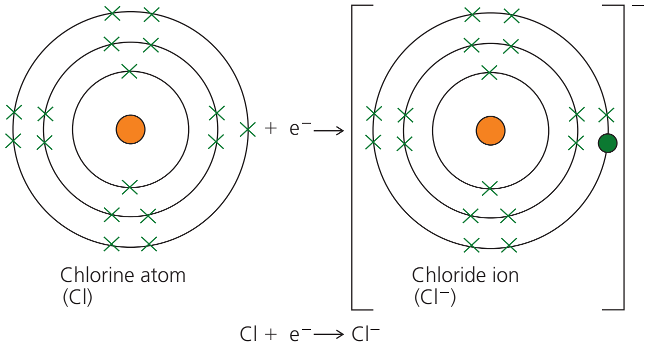
chlorine bromine iodine

You will notice that, in contrast to the elements of Groups I and II, the order of reactivity decreases on going down the group.

Table 9.5 shows the electronic configuration for chlorine and bromine. In each case, the outer electron shell contains seven electrons. When these elements react, they gain one electron per atom to gain the stable electronic configuration of a noble gas. You will learn more about the stable nature of these gases in the next section. For example, when chlorine reacts, it gains a single electron and forms a negative ion (Figure 9.14).

▼ **Table 9.5** Electronic configuration of chlorine and bromine

Element	Symbol	Proton number	Electronic configuration
Chlorine	Cl	17	2,8,7
Bromine	Br	35	2,8,18,7



▲ **Figure 9.14** A chlorine atom gains an electron to form a chloride ion

Chlorine is more reactive than bromine because the incoming electron is gained more easily by the smaller chlorine atom than in the larger bromine atom. It is gained more easily because there is a stronger attraction between the negative charge of the incoming electron and the positive charge of the nucleus. In the larger bromine atom, there are more occupied electron shells surrounding the nucleus, which lessen the attraction of the nucleus, and the electrons in these shells repel the incoming electron. This makes it harder for the bromine atom to gain the extra electron it needs to gain a stable electronic configuration. This is the reason the reactivity of the halogens decreases going down the group.

The halogens and their compounds are used in many different ways (Figure 9.15).

In the reaction of chlorine with potassium iodide, both Cl atoms in Cl_2 gain an electron from an iodide ion, I^-, thus forming two chloride ions, Cl^-. The iodine atoms formed by the loss of an electron combine to give an iodine molecule, I_2.

Oxidation:	$2I^-(aq)$ +	$Cl_2(g)$ →	$2Cl^-(aq)$ +	$I_2(aq)$
Number:	–1	0	–1	0
Colour:	colourless	green	colourless	orange-brown

The iodide ion has been oxidised because it has lost electrons. The oxidation number has increased. Chlorine has been reduced because it has gained electrons. The oxidation number has decreased.

Going further

Uses of the halogens

- Fluorine is used in the form of fluorides in drinking water and toothpaste because it reduces tooth decay by hardening the enamel on teeth.
- Chlorine is used to make PVC plastic as well as household bleaches. It is also used to kill bacteria and viruses in drinking water (Chapter 11, p. 174).
- Bromine is used to make disinfectants, medicines and fire retardants.
- Iodine is used in medicines and disinfectants, and also as a photographic chemical.

▲ **Figure 9.15** The halogens have many varied uses – fluoride in toothpaste to help reduce dental decay, iodine as a photographic chemical, chlorine in household bleach to kill bacteria, and bromine as a fire retardant

Practical skills

Halogen displacement reactions

For safe experiments/demonstrations which are related to this chapter, please refer to the *Cambridge IGCSE Chemistry Practical Skills Workbook*, which is also part of this series.

Safety

- Eye protection must be worn.

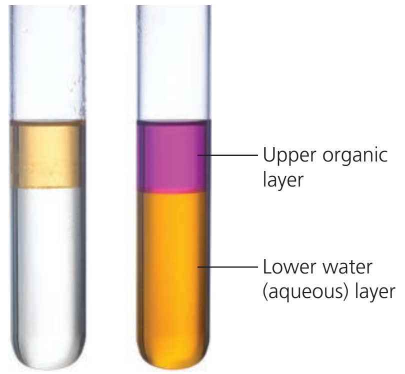
Displacement reactions can be used to determine the reactivity of the Group VII elements, the halogens. In this experiment, a student uses these reactions to determine an order of reactivity for iodine, chlorine and bromine, and predicts the reactivity of the other two halogens, fluorine and astatine.

If a more reactive halogen reacts with a compound of a less reactive halogen, the less reactive halogen will be displaced and it will form the halogen molecule, while the more reactive halogen becomes a halide ion.

To observe the presence of the different halogens in solution, the student used an organic solvent as the halogens produce more vivid colours in this solvent compared with water.

To show the colour of the halogens in the organic solvent solutions, the student separately placed three halogens in test tubes and added a small amount of the organic solvent. The tubes were then fitted with a rubber bung and shaken.

After shaking, two layers formed. The lower layer is the water (aqueous layer) and the upper layer is the organic layer.



▲ **Figure 9.16** The tube on the left shows the results of chlorine water being added to sodium bromide, and on the right the result of bromine water being added to sodium iodide

The results obtained are:

Halogen	Colour in water	Colour in organic solvent
Chlorine	Colourless	Pale green
Bromine	Orange	Orange
Iodine	Brown	Violet

The colours of the halogens in water can be used to identify if a reaction has occurred when a solution of a halogen, for example, chlorine water, is mixed with a solution of a sodium halide, for example, sodium iodide solution. The table below shows the results of six of these.

Sample data

		Chlorine water	Bromine water	Iodine water
Colour after shaking with sodium iodide solution		Brown	Brown	
Colour of each layer with the organic solvent added	Upper	Violet	Violet	
	Lower	Brown	Brown	
Colour after shaking with sodium bromide solution		Orange		Brown
Colour of each layer with the organic solvent added	Upper	Orange		Violet
	Lower	Pale orange		Brown
Colour after shaking with sodium chloride solution			Orange	Brown
Colour of each layer with the organic solvent added	Upper		Orange	Violet
	Lower		Pale orange	Brown

1 In how many of the six reactions has a displacement reaction occurred?
2 For each of these displacement reactions, write a balanced chemical equation.
3 What type of chemical change has happened to the halogen molecules in the displacement reactions?
4 State the order of reactivity of the three halogens used in the experiment.
5 What would be the colour of fluorine and astatine in water?
6 Which of the five halogen elements is the most reactive? Explain your answer.
7 Why were solutions of the halogens used, rather than the halogens themselves?

9.5 Group VIII – the noble gases

Helium, neon, argon, krypton, xenon and the radioactive element radon make up a most unusual group of non-metals, called the noble gases. They were all discovered after Mendeleev had published his Periodic Table. They were discovered between 1894 and 1900, mainly through the work of the British scientists Sir William Ramsay and Lord John William Strutt Rayleigh.

- They are colourless gases.
- They are monatomic gases – they exist as individual atoms, for example, He, Ne and Ar.
- They are very unreactive.

No compounds of helium, neon or argon have ever been found. However, more recently a number of compounds of xenon and krypton with fluorine and oxygen have been produced, for example, XeF_6.

These gases are chemically unreactive because they have electronic configurations which are stable and very difficult to change (Table 9.6). They are so stable that other elements attempt to attain these electronic configurations during chemical reactions (Chapter 3, p. 31, and p. 38). You have probably seen this in your study of the elements of Groups I, II and VII.

▼ **Table 9.6** Electronic configuration of helium, neon and argon

Element	Symbol	Proton number	Electronic configuration
Helium	He	2	2
Neon	Ne	10	2,8
Argon	Ar	18	2,8,8

Although unreactive, they have many uses. Argon, for example, is the gas used to fill light bulbs to prevent the tungsten filament reacting with air. Neon is used extensively in advertising signs and in lasers.

Test yourself

5 Write word and balanced chemical equations for the reactions between:
   a bromine and potassium iodide solution
   b bromine and potassium chloride solution.
   If no reaction will take place, write 'no reaction' and explain why.
   c Using the information on pp. 141–142, predict the properties of the element astatine related to its melting point, density and physical state at room temperature. Predict how astatine would react with sodium bromide solution.

9.6 Transition elements

This block of metals includes many you will be familiar with, for example, copper, iron, nickel, zinc and chromium (Figure 9.17).

- They are less reactive metals.
- They form a range of brightly coloured compounds (Figure 9.18).
- They are harder and stronger than the metals in Groups I and II.
- They have much higher densities than the metals in Groups I and II.
- They have high melting points (except for mercury, which is a liquid at room temperature).
- They are good conductors of heat and electricity.
- They show catalytic activity (Chapter 7, p. 106) as elements and compounds. For example, iron is used in the industrial production of ammonia gas (Haber process, Chapter 7, p. 109).
- They do not react (corrode) so quickly with oxygen and/or water.

a Copper is used in many situations which involve good heat and electrical conduction. It is also used in medallions and bracelets

b These gates are made of iron. Iron can easily be moulded into different shapes

c Monel is an alloy of nickel and copper. It is extremely resistant to corrosion, even that caused by sea water

d This bucket has been coated with zinc to prevent the steel of the bucket corroding

e The alloy stainless steel contains a high proportion of chromium, which makes it corrosion resistant

▲ **Figure 9.17** Everyday uses of transition elements and their compounds. They are often known as the 'everyday metals'

a Some solutions of coloured transition element compounds

b The coloured compounds of transition elements can be seen in these pottery glazes

▲ **Figure 9.18**

- They form simple ions with variable oxidation numbers. (For a discussion of oxidation number, see Chapter 3, p. 35.) For example, copper forms Cu^+ (Cu(I)) and Cu^{2+} (Cu(II)), in compounds such as Cu_2O and $CuSO_4$, and iron forms Fe^{2+} (Fe(II)) and Fe^{3+} (Fe(III)), in compounds such as $FeSO_4$ and $FeCl_3$.
- They form more complicated ions with high oxidation numbers. For example, chromium forms the dichromate(VI) ion, $Cr_2O_7^{2-}$, which contains chromium with a +6 oxidation state (Cr(VI)) and manganese forms the manganate(VII) ion, MnO_4^-, which contains manganese with a +7 oxidation number (Mn(VII)).

Test yourself

6 Look at the photographs in Figure 9.17 and state which key properties are important when considering the particular use of the metal.

7 Which groups in the Periodic Table contain:
   - **a** only metals?
   - **b** only non-metals?
   - **c** both metals and non-metals?

9.7 The position of hydrogen

Hydrogen is often placed by itself in the Periodic Table. This is because the properties of hydrogen are unique. However, useful comparisons can be made with the other elements. It is often shown at the top of either Group I or Group VII, but it cannot fit easily into the trends shown by either group; see Table 9.7.

▼ **Table 9.7** Comparison of hydrogen with lithium and fluorine

Lithium	Hydrogen	Fluorine
Solid	Gas	Gas
Forms a positive ion	Forms positive or negative ions	Forms a negative ion
1 electron in outer electron shell	1 electron in outer electron shell	1 electron short of a full outer electron shell
Loses 1 electron to form a noble gas configuration	Needs 1 electron to form a noble gas configuration	Needs 1 electron to form a noble gas configuration

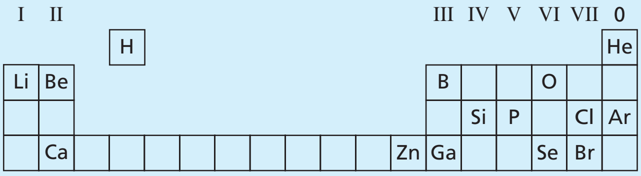
Revision checklist

After studying Chapter 9 you should be able to:

- ✓ Use the Periodic Table to help you to predict the properties of elements.
- ✓ Use the Periodic Table to help you to give the charges on ions and to write chemical and ionic equations.
- ✓ Describe the trends in the properties and reactions of the Group I and Group VII elements.
- ✓ Describe the Group VIII elements as unreactive, monoatomic gases and explain why they are unreactive.
- ✓ Identify trends in the different groups if you are given information.
- ✓ Describe the physical appearance of the Group I and VII elements.
- ✓ Describe halogen-halide displacement reactions.
- ✓ Give the properties of the transition metals.
- ✓ Understand that transition metal ions have variable oxidation numbers.

Exam-style questions

1 The diagram below shows part of the Periodic Table.



Using **only** the symbols of the elements shown above, give the symbol for an element which:

a is a pale yellow-green coloured toxic gas [1]
b is stored under oil [1]
c has five electrons in its outer electron energy shell [1]
d is the most reactive Group II element [1]
e is the most reactive halogen [1]
f is the only liquid shown [1]
g is a transition element [1]
h is a gas with two electrons in its outer shell. [1]

2 Three members of the halogens are $^{35.5}_{17}Cl$, $^{80}_{35}Br$ and $^{127}_{53}I$.

a **i** Give the electronic configuration of an atom of chlorine. [1]
ii Explain why the relative atomic mass of chlorine is not a whole number. [1]
iii Identify how many protons there are in an atom of bromine. [1]
iv Identify how many neutrons there are in an atom of iodine. [1]
v Explain the order of reactivity of these elements. [5]

b When potassium is allowed to burn in a gas jar of chlorine, in a fume cupboard, clouds of white smoke are produced.
i Explain why this reaction is carried out in a fume cupboard. [1]
ii State what the white smoke consists of. [1]
iii Give a word and balanced chemical equation for this reaction. [2]
iv Describe what you would expect to see when potassium is allowed to burn safely in a gas jar of bromine vapour. Write a word and balanced chemical equation for this reaction. [3]

3 'By using displacement reactions, it is possible to deduce the order of reactivity of the halogens.' Discuss this statement with reference to the elements bromine, iodine and chlorine only. [4]

4 Use the information given in the table below to answer the questions concerning the elements **Q**, **R**, **S**, **T** and **X**.

Element	Proton number	Mass number	Electronic configuration
Q	3	7	2,1
R	20	40	2,8,8,2
S	18	40	2,8,8
T	8	18	2,6
X	19	39	2,8,8,1

a Identify the element that has 22 neutrons in each atom. [1]
b Identify the element that is a noble gas. [1]
c Identify the two elements that form ions with the same electronic configuration as argon. [2]
d Identify the two elements that are in the same group of the Periodic Table and which group this is. [2]
e Place the elements in the table into the periods in which they belong. [3]

f Identify the most reactive metal element in the table. [1]

g Identify which of the elements is calcium. [1]

5 **a** Consider the chemical properties and physical properties of the halogens chlorine, bromine and iodine. Using these properties, predict the following about the other two halogens, fluorine and astatine. [2]

Property	Fluorine	Astatine
State at room temperature and pressure		
Colour		
Reactivity with sodium metal		

b **i** Give a word equation for the reaction of chlorine gas with sodium bromide solution. [1]

ii Give a balanced chemical equation for the reaction, with state symbols. [2]

iii Give an ionic equation for the reaction, with state symbols. [2]

6 Some of the most important metals we use are found in the transition element section of the Periodic Table. One of these elements is copper. Sodium, a Group I metal, has very different properties from those of copper. Complete the table below to show their differences. [4]

	Transition element, e.g. copper	**Group I metal, e.g. sodium**
Hardness (hard/soft)		
Reactivity		
Density (high/low)		
Variable oxidation states		