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DYNAMIC LEARNING
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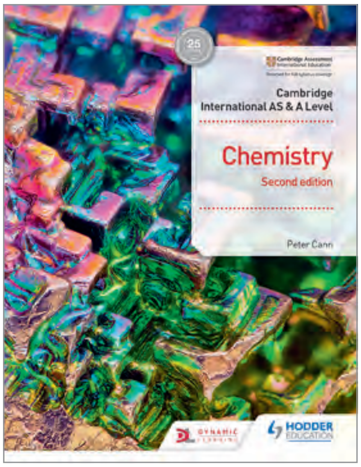

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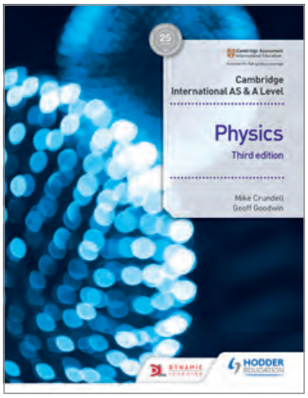

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If we can help with questions, and to find out more, please contact us at **international.sales@hoddereducation.com**

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10

Periodicity

A modern form of the periodic table is studied in this topic. We shall see how the periodic table can be used to show patterns in the properties of some of the chemical elements it contains. The changes in properties in crossing Period 3 (from sodium to argon) are related to the changes in the elements' atomic structures.

Learning Outcomes

By the end of this topic, you will be able to:

- describe qualitatively (and indicate the periodicity in) the variations in atomic radius, ionic radius, melting point and electrical conductivity of the elements in Period 3
- explain the variation in melting point and electrical conductivity in terms of the structure and bonding of the elements In Period 3
- describe, and write equations for, the reactions of the elements in Period 3 with oxygen (to give Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2), chlorine (to give $NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5) and water (Na and Mg only)
- state and explain the variation in the oxidation number of the oxides (Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2 and SO_3 only) and chlorides ($NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5 only) in terms of their outer shell (valence shell) electrons
- describe and write equations for the reactions, if any, of the oxides Na_2O, MgO, Al_2O_3, SiO_2, P_4O_{10}, SO_2 and SO_3 with water including the likely pHs of the solutions obtained
- describe, explain, and write equations for, the acid/base behaviour of the oxides Na_2O, MgO, Al_2O_3, P_4O_{10}, SO_2 and SO_3 and the hydroxides $NaOH$, $Mg(OH)_2$, $Al(OH)_3$ including, where relevant, amphoteric behaviour in reactions with acids and bases (sodium hydroxide only)
- describe, explain, and write equations for, the reactions of the chlorides $NaCl$, $MgCl_2$, $AlCl_3$, $SiCl_4$, PCl_5 with water including the likely pHs of the solutions obtained
- explain the variations and trends in 9.2.2, .3, .4 and .5 in terms of bonding and electronegativity
- suggest the types of chemical bonding present in chlorides and oxides from observations of their chemical and physical properties
- predict the characteristic properties of an element in a given group by using knowledge of chemical periodicity
- deduce the nature, possible position in the periodic table and unknown elements from given information about physical and chemical properties.

10.1 Background

In the nineteenth century, when about 50 elements had been discovered, attempts were made to find patterns in their properties. The following three attempts were particularly important:

- In 1829, Johann Döbereiner pointed out that some similar elements formed groups of three (which he called ‘triads’). The middle element showed properties intermediate between the other two. Examples of these triads included lithium, sodium and potassium; calcium, strontium and barium; chlorine, bromine and iodine.
- In 1864, John Newlands showed that if the elements are arranged in order of their relative atomic masses, many show properties similar to the element that is eight places further on. He called this the Law of Octaves, after the eight white notes on a piano which make up an octave. Examples include sodium and potassium; magnesium and calcium; boron and aluminium; oxygen and sulfur; nitrogen and phosphorus; fluorine and chlorine.
- The first versions of the periodic table were independently produced by Julius Lothar Meyer and by Dmitri Mendeleev in 1870. Lothar Meyer was the first to put his ideas down on paper, but publication of his manuscript was delayed for two years, and so Mendeleev is often given sole credit.

▲ Figure 10.1 Dmitri Mendeleev (1834–1907) collected information about the elements and wrote them down on cards, which he arranged according to relative atomic mass and to the properties of the elements. Along with Julius Lothar Meyer, he formed the periodic table we use today

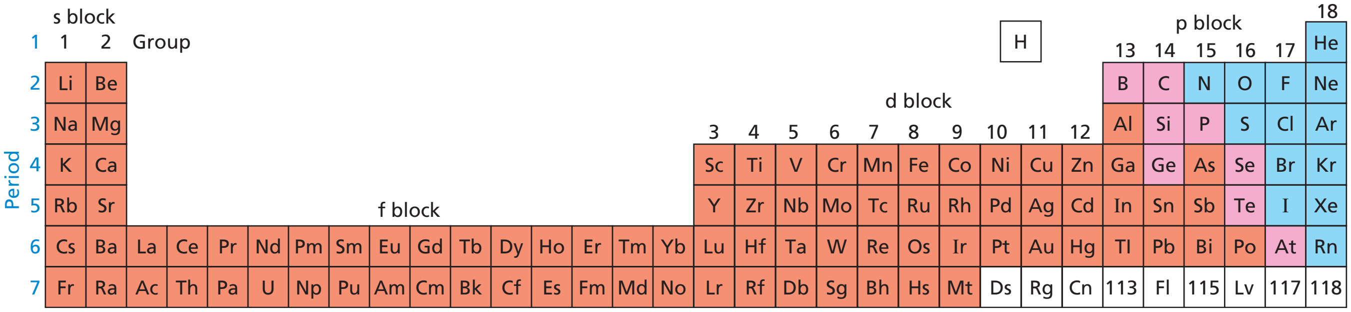
Two features of Mendeleev’s classification were particularly significant.

- He left gaps in his table for elements that were unknown at that time. He successfully predicted the existence of the elements scandium, gallium and germanium, all of which were discovered a few years later. He used Sanskrit prefixes to identify his ‘missing elements’; for example, his name for germanium was eka-silicon, meaning ‘one place below silicon’ in his table.
- He reversed the orders of cobalt and nickel, and tellurium and iodine, from the order shown by their relative atomic masses. Subsequent work showed that this reversal was correct. He had put them in the order of their atomic numbers; the fact that their relative atomic masses were in the wrong order was because of the proportions of isotopes in the elements (see Topic 2).

10.2 A modern form of the periodic table

Groups and periods

A modern form of the periodic table is shown in Figure 10.2.



	s block		f block														d block										p block					
Period	1	2 Group															3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1																								H								He
2	Li	Be																									B	C	N	O	F	Ne
3	Na	Mg																									Al	Si	P	S	Cl	Ar
4	K	Ca															Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
5	Rb	Sr															Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
6	Cs	Ba	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn
7	Fr	Ra	Ac	Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	113	Fl	115	Lv	117	118

▲ Figure 10.2 A modern form of the periodic table. The elements shown in orange are good conductors of electricity, and the ones shown in blue are poor conductors. The pink shading indicates that the principal form of the element is a semiconductor. As only a few atoms of the elements at the end of the periodic table have been made, these have been left white

Elements placed under each other in a column are in the same **group** and show many similarities in their physical and chemical properties. Elements are also arranged in rows or **periods** across the table. The elements in a period have different physical and chemical properties, but trends become apparent in these properties as we move across a period.

The s, p, d and f blocks

The first two groups (Groups 1 and 2) form the **s block**. These have the outer electronic configuration ns^1 and ns^2, respectively, where n is the number of the shell. The last six groups (Groups 13 to 18) form the **p block** in which the p subshell is being progressively filled. These elements have the outer electronic configuration $ns^2\ np^1$ to $ns^2\ np^6$.

In between the s and p blocks is the **d block**, in which the d subshell is being progressively filled. Our study of the d block is largely restricted to the elements scandium to zinc, and these elements have outer electronic configurations $4s^2\ 3d^1$ to $4s^2\ 3d^{10}$.

About a quarter of the known elements belong to the **f block**. The first row, the 4f, includes all the elements from cerium (Ce) to lutetium (Lu) inclusive. In this block, the f subshell in the fourth principal shell is being progressively filled. In the past, these elements were called the 'rare earths'. This was a misnomer, as they are fairly abundant in the Earth's crust – for some their abundance is comparable to that of lead. However, it is true that they are mostly quite thinly spread, so mining them is expensive. They are now called the **lanthanides** as their properties are similar to those of the element lanthanum (La), which precedes them in the periodic table. Although they are not particularly rare, the lanthanides are difficult to purify from each other because their chemical properties are nearly identical. The elements of the second row, the 5f, are called the **actinides**. All the actinides are radioactive and most have to be made artificially.

10.3 Periodic trends in the physical properties of the elements of the third period (sodium to argon)

Appearance

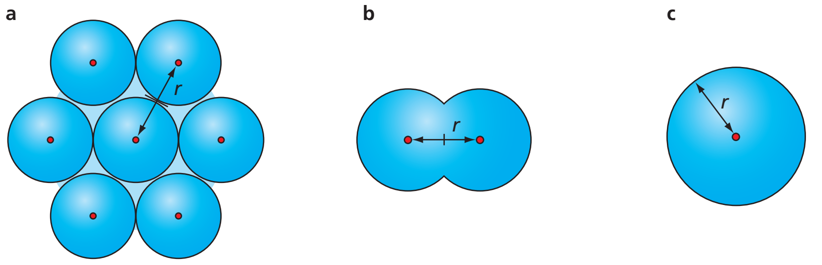
The elements on the left of the periodic table have low values of ionisation energy and electronegativity, and so they show the properties associated with metallic bonding; for example, they are shiny and conduct electricity. In the middle of the periodic table, elements with higher values of ionisation energy and electronegativity are semiconductors: they have a dull shine to them and are poor conductors of electricity (typically 10^{-12} times that of a metal). The elements at the right of the periodic table, with the highest values of ionisation energy and electronegativity, are dull in appearance and are such poor conductors that they are used as electrical insulators (their conductivities are virtually zero, being typically 10^{-18} times that of a metal).

Property	Na	Mg	Al	Si	P	S	Cl	Ar
Formula of common ion	Na^+	Mg^{2+}	Al^{3+}	(Si^{4+})	P^{3-}	S^{2-}	Cl^-	
Atomic radius/nm	0.186	0.160	0.143	0.117	0.110	0.104	0.099	(0.188)
Ionic radius/nm	0.095	0.065	0.050	–	0.212	0.184	0.181	
Melting point/°C	98	649	660	1410	44	113	−101	−189
Electrical conductivity/S cm^{-1}	2.1×10^5	2.3×10^5	3.8×10^5	2.5×10^{-6}	1.0×10^{-11}	5.0×10^{-18}	0	0
First ionisation energy/kJ mol^{-1}	494	736	577	786	1060	1000	1260	1520
Electronegativity	0.9	1.2	1.5	1.8	2.1	2.5	3.0	–
Common oxidation state(s)	+1	+2	+3	+4	−3, +3, +5	−2, +4, +6	−1, +1, +5, +7	–

▲ Table 10.1 Some properties of the elements in the third period

Interatomic and interionic radii

The term 'atomic radius' does not mean the same thing for every element. Although we can fairly easily measure the distance between two adjacent nuclei in an element (see Figure 10.3), if we are to use those distances to compare one element with another, we must be aware that the interatomic bonding might differ in the two elements. For example, tables of interatomic radii include three values for sodium: its covalent radius is 0.154 nm; its metallic radius is 0.186 nm; and its van der Waals radius is 0.227 nm. The only radius listed for argon is its van der Waals radius (0.188 nm), since it does not form compounds. If we are to compare the sizes of the atoms in a sensible way, we need to compare radii of the same type. Comparing the van der Waals radii of the two elements shows that the sodium atom (0.227 nm) is larger than the argon atom (0.188 nm), as we might expect, since the nuclear attraction for the outer electron shell increases with proton number across the group.



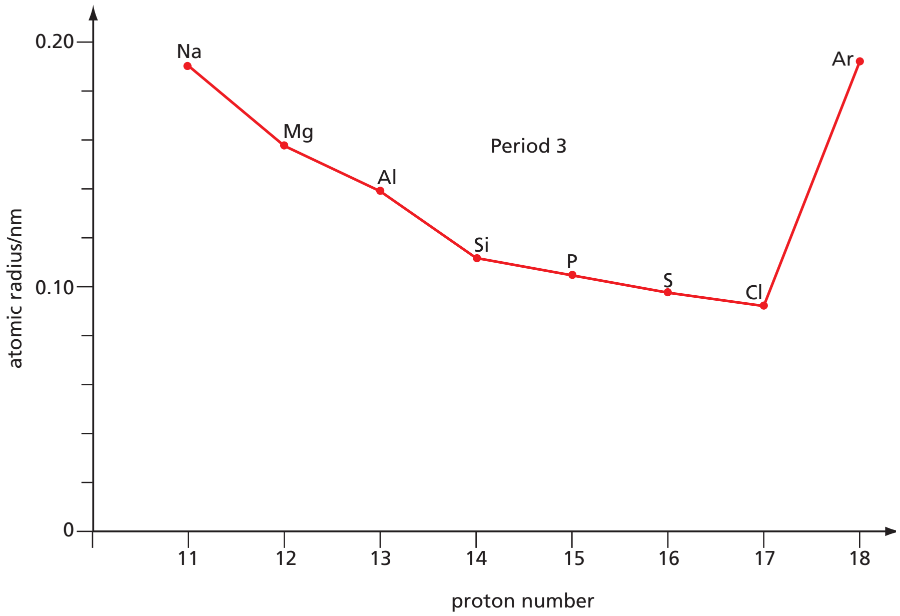
on the left – metallic radius

on the right – covalent radius

for Group 18 –
van der Waals' radius

▲ Figure 10.3 The atomic radius is half the internuclear distance

In crossing the third period of the periodic table, the proton number increases from 11 to 18. This means that the 3s and 3p outer electrons become more firmly attracted to the nucleus and, as a result, their distance away from the nucleus reduces. This is the reason why the atomic radii decrease from sodium to chlorine (Figure 10.4 overleaf). Argon does not have an atomic radius as it does not form a bond to itself. The large value of 0.188 nm listed in the table is the van der Waals radius, that is half the distance when the atoms touch each other in the solid state.



▲ **Figure 10.4 Atomic radius and proton number**

Sodium, magnesium and aluminium are metals and form Na^+,Mg^{2+} and Al^{3+} ions. These ions all have the electronic structure $1s^2\ 2s^2\ 2p^6$. As the proton number increases from 11 with sodium to 13 with aluminium, the attraction for the outer electrons increases and they become drawn in closer to the nucleus. This means that the ionic radii get smaller (see Table 10.1).

Phosphorus, sulfur and chlorine are non-metals which form ions by filling their outer shells. All have the electronic structure $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$. They are all larger than the ions of the metals in Period 3 because they have an extra shell of electrons. With an increase in proton number, from 15 in phosphorus to 17 in chlorine, the attraction for the outer electrons increases and so the ionic radius decreases.

Now try this 10.1

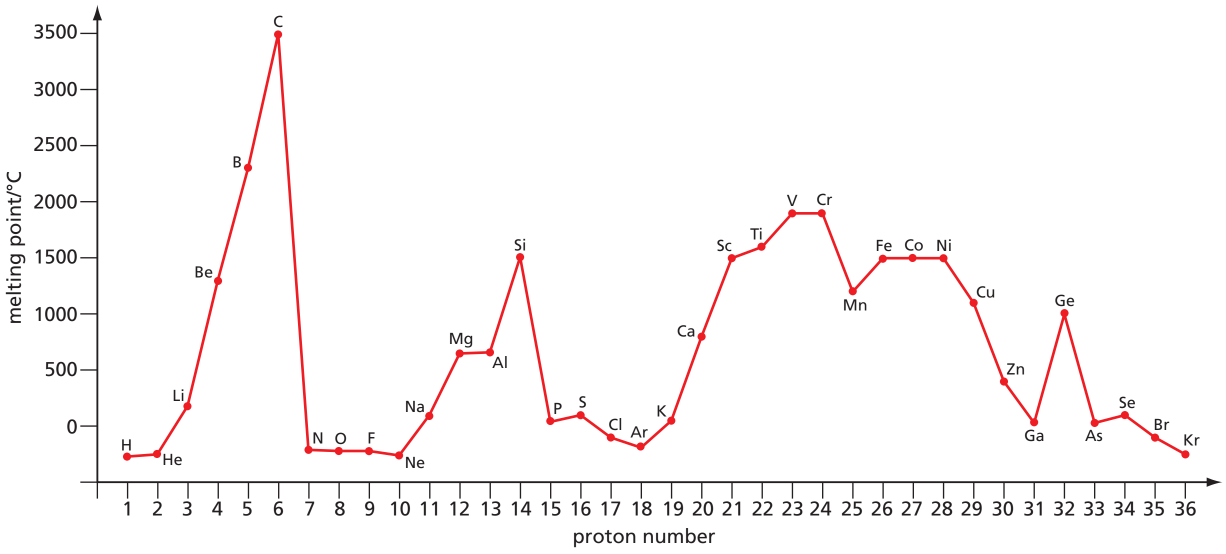
All the common compounds of silicon are covalent, so any ionic radii that are listed in data tables tend to be estimates. In theory, silicon could gain a noble gas electronic configuration by losing four electrons to give Si^{4+}, or by gaining four electrons to give Si^{4-}.

Predict values for the radii of these two ions, explaining your reasoning.

Melting points

The melting point of an element that has a giant structure is high because interatomic bonds must be broken for melting to take place. The atomic radii become smaller from sodium to silicon, and so the bonding becomes stronger and the melting points become higher, as the bonding electrons are closer to the adjacent nuclei.

Phosphorus, sulfur and chlorine all form small covalent molecules, P_4, S_8 and Cl_2, respectively. When these substances melt, it is only necessary to break weak intermolecular bonds and not strong interatomic attractions. The melting points decrease in the order sulfur > phosphorus > chlorine (see Figure 10.5): the intermolecular forces (mainly id–id) become weaker as the molecules become smaller (see Section 3.17). The melting point of argon is very low as the id–id attraction between the argon atoms is very small (Figure 10.5).



▲ **Figure 10.5 Melting point and proton number**

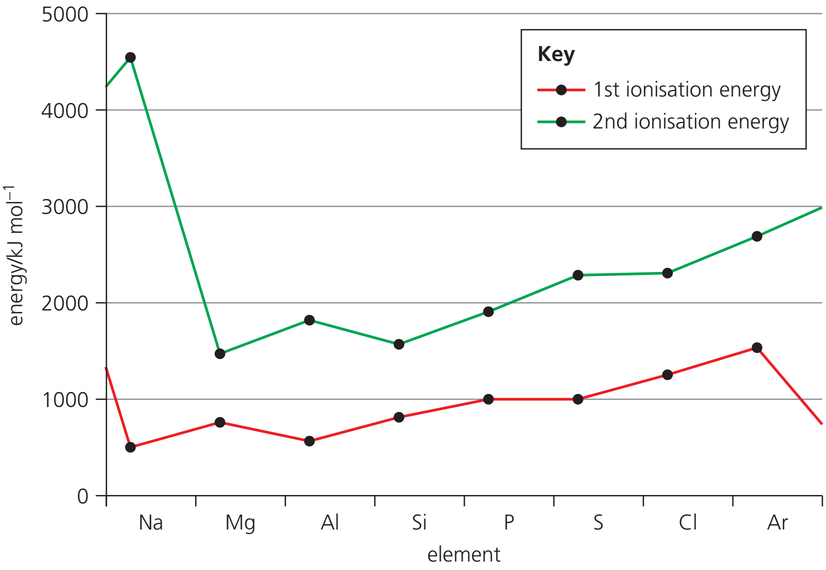
Electrical conductivity

Sodium, magnesium and aluminium are metals. They have delocalised electrons that are free to move through the lattice of cations (see Section 4.11). Silicon is a semiconductor. The other elements form covalent bonds with no free electrons and so are insulators with almost no electrical conductivity.

Ionisation energies

The first ionisation energies generally increase from sodium to argon as the proton number increases. In two instances, this increase does not take place: between magnesium and aluminium, and between phosphorus and sulfur (see Figure 10.6).

The second ionisation energies show a similar trend, but displaced by one atomic number unit. The second ionisation energy for sodium is very high due to the removal of an electron from an inner shell, nearer to the nucleus.



▲ **Figure 10.6 First and second ionisation energy against proton number (z), from z = 11 to z = 18**

Now try this 10.2

1 Suggest why:
  - **a** the ionisation energy of aluminium is less than that of magnesium
  - **b** the ionisation energy of sulfur is less than that of phosphorus.

Hint: You might need to return to Section 2.11 to refresh your memory.

Now try this 10.3

1 The two elements silicon and germanium have higher melting points than their neighbours (see Figure 10.5).
  - **a** Suggest the bonding and structure in germanium.
  - **b** Would you expect germanium to be a conductor, a semiconductor or an insulator?

2 Suggest why the melting point of selenium (proton number 34) is higher than that of arsenic (proton number 33).

10.4 Periodic trends in the chemical properties of the elements of the third period (sodium to sulfur)

Reactions of the elements with oxygen and chlorine

(N.B. the reaction of sulfur with chlorine, and the properties of the various sulfur chlorides, are not part of the syllabus.)

All the elements except chlorine and argon react directly and exothermically with oxygen and with chlorine, but they often need strong initial heating to overcome the energy barrier to reaction.

Sodium burns with a yellow flame; magnesium and phosphorus burn with bright white flames; and sulfur burns with a blue flame. The reactions occurring when chlorine is passed over the heated elements are slightly less vigorous: flames are only seen with sodium and magnesium.

with oxygen	**with chlorine**
$2Na(s) + \frac{1}{2}O_2(g) \rightarrow Na_2O(s)$	$Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$
$Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$	$Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$
$2Al(s) + 1\frac{1}{2}O_2(g) \rightarrow Al_2O_3(s)$	$2Al(s) + 3Cl_2(g) \rightarrow Al_2Cl_6(s)$
$Si(s) + O_2(g) \rightarrow SiO_2(s)$	$Si(s) + 2Cl_2(g) \rightarrow SiCl_4(l)$
$P_4(s) + 3O_2(g) \xrightarrow{\text{limited oxygen}} P_4O_6(s)$	$P_4(s) + 6Cl_2(g) \xrightarrow{\text{limited chlorine}} 4PCl_3(l)$
$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$	$P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s)$
$S(s) + O_2(g) \rightarrow SO_2(g)$	

$$SO_2(g) + \frac{1}{2}O_2(g) \xrightarrow{V_2O_5\text{ catalyst}} \underset{\text{sulfur trioxide}}{SO_3(g)}$$

The oxidation numbers of the elements in their oxides and chlorides are usually equal to the number of electrons in the $n = 3$ shell. For the ionic compounds, Na_2O, NaCl, $MgCl_2$, MgO and Al_2O_3, the metals have formed Na^+, Mg^{2+} and Al^{3+} ions by loss of all the outer electrons. The other oxides and chlorides are covalent. In these covalent compounds the oxidation number of the element can be found by giving chlorine an oxidation number of (–1) and oxygen an oxidation number of (–2). With the exception of SO_2 and PCl_3, the oxidation number of the element is the same as the number of electrons in its outer shell. This is because all of the electrons are used to make bonds. In SO_2, two of the electrons of the sulfur form a lone pair and only four are used in bonding to oxygen. This then gives sulfur an oxidation number of (+4), two fewer than the six electrons in its outer shell.

Electronegativity increases across the period (see Table 10.1 for some values; also see Section 3.10 for an explanation of the origin of electronegativity). The difference in electronegativity between the Period 3 element and oxygen or chlorine is high at the left-hand side and gets less and less going across the period. For Na_2O, NaCl, $MgCl_2$, MgO and Al_2O_3, the electronegativity difference is very high and so these compounds are ionic; they have high melting points and in the molten state they conduct electricity. The electronegativity differences between Si, P and S and oxygen, and between Al, Si and P and chlorine, are quite low, which would predict covalent bonding. The low melting points of the oxides of phosphorus and sulfur, and the chlorides of silicon and phosphorus suggest they exist as small covalent molecules. SiO_2 exists as a high melting-point, giant covalent lattice.

SO_2, SO_3 and P_4O_{10} contain small covalent molecules. Al_2Cl_6, $SiCl_4$ and PCl_5 are also covalent.

This is shown by their low melting points:

- SO_2 is a gas, SO_3 is a low melting-point solid and P_4O_{10} sublimes when heated gently.
- Al_2Cl_6 readily sublimes, $SiCl_4$ is a liquid and PCl_5 decomposes when heated above its low melting point to give PCl_3 which is a liquid.

	Sodium	**Magnesium**	**Aluminium**	**Silicon**	**Phosphorus**	**Sulfur**
Melting point of oxide/°C	1275	2852	2072	1710	42 (P_4O_{10})	17 (SO_3)
Melting point of chloride/°C	801	714	190	–70	–94 (PCl_3)	–78 (SCl_2)

▲ Table 10.2 Melting points of the oxides and chlorides of Period 3 elements

WORKED EXAMPLE 10A

Element X forms a chloride with melting point of –49 °C and an oxide with melting point 1115 °C. Neither compound conducts electricity when molten.

Suggest which periodic table group element X is in, the type of bonding in its oxide and chloride, and the formulae of its oxide and chloride. Explain your answer.

Answer

Since neither compound conducts electricity when molten, they must both be covalent compounds. The low melting point of the chloride suggests a simple molecular structure, whereas the high melting point of the oxide suggests a giant covalent structure. This is typical of elements in the centre of the periodic table, similar to silicon in Period 3, Group 14. Therefore, element X is likely to be in Group 14. So, the formulae would be XCl_4 and XO_2.

(In fact, X = Ge)

Now try this 10.4

1 Use the properties of the following compounds to predict their type of bonding. Then suggest the bonding in, and one other property of, the second compound.

a Element Y forms a chloride with melting point −10 °C. Its oxide melts at 311 °C.

b Element Z forms an oxide with melting point 1918 °C. Its chloride dissolves in water to give a solution of pH 7.

Reactions of the elements with water

Only sodium and magnesium react with water. Sodium reacts vigorously with water in the cold but, in order to result in an appreciable reaction, magnesium needs to be red hot to react with steam:

$$Na(s) + H_2O(l) \rightarrow NaOH(aq) + \tfrac{1}{2}H_2(g)$$

$$Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$$

Magnesium *will* react with cold water, but only very slowly: it takes several hours to collect a test tube of hydrogen.

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

Reactions of the oxides with water, acids or alkalis

When an oxide of an element M reacts with water, a M-O-H bond is formed. This bond may ionise in two ways.

$$M\text{-}O\text{-}H \rightarrow M^+ + OH^-$$

or

$$M\text{-}O\text{-}H \rightarrow MO^- + H^+$$

In the first example, the hydroxide is behaving as a base, in the second as an acid. If the element M has a low electronegativity (<1.5), the first reaction takes place and the hydroxide is basic. If M has a higher electronegativity (>1.8), the second reaction takes place and the hydroxide behaves as an acid. If M has an electronegativity between 1.5 and 1.8, both reactions are possible and the hydroxide is amphoteric.

Sodium oxide (electronegativity 0.7) easily dissolves in water to give an alkaline solution:

$$Na_2O(s) + H_2O(l) \rightarrow 2Na^+OH^-(aq)$$

pH = 13–14

The hydroxide reacts with acids to form neutral salts:

$$NaOH(aq) + HNO_3(aq) \rightarrow Na^+NO_3^-(aq) + H_2O(l)$$

Magnesium oxide (electronegativity 1.2) dissolves slightly in water to give a solution that is just alkaline:

$$Mg(s) + H_2O(l) \rightarrow Mg^{2+}(OH^-)_2(aq)$$

pH = 10–11

The oxide reacts readily with acids to form salts, most of which are soluble:

$$MgO(s) + H_2SO_4(aq) \rightarrow Mg^{2+}SO_4^{2-}(aq) + H_2O(l)$$

Aluminium oxide does not dissolve in water, but its hydroxide (electronegativity 1.5) is amphoteric (it dissolves in both acids and alkalis):

$$Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$$

$$Al(OH)_3(s) + OH^-(aq) \rightarrow Al(OH)_4^-(aq)$$

SiO_2 is a weakly acidic oxide and is the starting material for glasses and ceramics because, when molten, it reacts with bases such as sodium carbonate and calcium carbonate.

P_4O_{10} and SO_2 (electronegativities 2.1 and 2.5) react with water to form weak acids, H_3PO_4 and H_2SO_3, while SO_3 forms the strong acid, H_2SO_4:

$$P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$$

pH = 2–3

$$SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$$

pH = 1–2

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

pH = 0–1

These non-metallic oxides, or the acids derived from them, react with alkalis to give neutral or acidic salts, depending on the stoichiometry:

$H_3PO_4(aq) + 2NaOH(aq) \rightarrow 2H_2O(l) + Na_2HPO_4(aq)$	(an acidic salt)
$SO_2(g) + Ca(OH)_2(aq) \rightarrow H_2O(l) + CaSO_3(aq)$	(a neutral salt)
$H_2SO_4(aq) + KOH(aq) \rightarrow H_2O(l) + KHSO_4(aq)$	(an acidic salt)
$H_2SO_4(aq) + 2KOH(aq) \rightarrow 2H_2O(l) + K_2SO_4(aq)$	(a neutral salt)

SUMMARY

- The elements of Period 3 show a general trend in their properties from left to right:
  - as nuclear charge increases, so does electronegativity
  - as electronegativity increases, they show a trend from metals to non-metals
  - as nuclear charge increases, their atomic radii decrease, and their first ionisation energies show a general increase (with exceptions after Mg and P)
  - their melting points reflect their structures – giant structures on the left have high melting points, simple molecular structures on the right have low melting points
  - they show a trend from electrical conductors to semiconductors to insulators.
- The compounds of Period 3 elements show the following trends:
  - the oxidation number of the element increases from +1 to +6 in their oxides (and chlorides up to PCl_5)
  - their oxides show a trend from basic to amphoteric to acidic
  - the structures of their oxides and chlorides show a trend from giant ionic to simple covalent
  - the pH of the solutions resulting from the reaction between their chlorides and water show a trend from neutral to strongly acidic.

Key reactions you should know

- Elements with oxygen:
  - $2Na(s) + \frac{1}{2}O_2(g) \rightarrow Na_2O(s)$
  - $Mg(s) + \frac{1}{2}O_2(g) \rightarrow MgO(s)$
  - $2Al(s) + 1\frac{1}{2}O_2(g) \rightarrow Al_2O_3(s)$
  - $Si(s) + O_2(g) \rightarrow SiO_2(s)$
  - $P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$
  - $S(s) + O_2(g) \rightarrow SO_2(g)$ (limited supply of oxygen)
  - $S(s) + 1\frac{1}{2}O_2(g) \rightarrow SO_3(g)$ (an excess of oxygen)
  - $SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g)$ (400 °C, V_2O_5 catalyst, pressure > 1 atm)
- Elements with chlorine:
  - $Na(s) + \frac{1}{2}Cl_2(g) \rightarrow NaCl(s)$
  - $Mg(s) + Cl_2(g) \rightarrow MgCl_2(s)$ (Cl_2 could be replaced by Br_2 or I_2)
  - $2Al(s) + 3Cl_2(g) \rightarrow Al_2Cl_6(s)$
  - $Si(s) + 2Cl_2(g) \rightarrow SiCl_4(l)$
  - $P_4(s) + 10Cl_2(g) \rightarrow 4PCl_5(s)$
- Elements with water:
  - $Na(s) + H_2O(l) \rightarrow NaOH(aq) + \frac{1}{2}H_2(g)$
  - $Mg(s) + H_2O(g) \rightarrow MgO(s) + H_2(g)$
- Oxide/hydroxide with water/acid:
  - $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$
  - $NaOH(aq) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l)$
  - $MgO(g) + 2H^+(aq) \rightarrow Mg^{2}+(aq) + H_2O(l)$

- $Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$
- $Al(OH)_3(s) + 3H^+(aq) \rightarrow Al^{3+}(aq) + 3H_2O(l)$
  (Also $Al(OH)_3(s) + OH^-(aq) \rightarrow Al(OH)_4^-(aq)$)
- $P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$
- $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$
- $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

- Chlorides with water:
  - $NaCl(s) + aq \rightarrow Na^+(aq) + Cl^-(aq)$ (just dissolves - no reaction)
  - $MgCl_2(s) + aq \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$
  - $Mg^{2+}(aq)$ is $[Mg(H_2O)^6]^{2+}(aq)$
    $[Mg(H_2O)_6]^{2+}(aq) + H_2O(l) \rightleftharpoons [Mg(H_2O)_5(OH)]^+(aq) + H_3O^+(aq)$
  - $AlCl_3(s) + 6H_2O(l) \rightleftharpoons [Al(H_2O)_6]^{3+}(aq) + 3Cl^-(aq)$
  - $[Al(H_2O)_6]^{3+}(aq) + H_2O(l) \rightleftharpoons [Al(H_2O)_5(OH)]^{2+}(aq) + H_3O^+(aq)$
  - $SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl(g)$
  - $PCl_5(s) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(g)$

Practice examination questions

							H										He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr

1 This question refers to the elements shown in the portion of the periodic table given above.

a From this table, identify in **each** case **one** element that has the property described. Give the **symbol** of the element in each case.

i The element that has a molecule which contains exactly eight atoms.
ii The element that forms the largest cation.
iii An element that floats on water and reacts with it.
iv An element that reacts with water to give a solution that can behave as an oxidising agent.
v An element whose nitrate gives a brown gas on thermal decomposition. **[5]**

b **i** Give the formula of the oxide of the most electronegative element.
ii Several of these elements form more than one acidic oxide. Give the formulae of two such oxides formed by the same element. **[3]**

The formulae and melting points of the fluorides of the elements in Period 3, Na to Cl, are given in the table.

Formula of fluoride	NaF	MgF_2	AlF_3	SiF_4	PF_5	SF_6	ClF_5
m.p./K	1268	990	1017	183	189	223	170

c **i** Suggest the formulae of two fluorides that could possibly be ionic.
ii What is the shape of the SF_6 molecule?
iii In the sequence of fluorides above, the oxidation number of the elements increases from NaF to SF_6 and then falls at ClF_5.

Attempts to make ClF_7 have failed but IF_7 has been prepared.

Suggest an explanation for the existence of IF_7 and for the non-existence of CIF_7. **[4]**

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2 Chlorine is very reactive and will form compounds by direct combination with many elements.

a Describe what you would see when chlorine is passed over separate heated samples of sodium and phosphorus. In each case write an equation for the reaction. **[4]**

b Magnesium chloride, $MgCl_2$, and silicon tetrachloride, $SiCl_4$, each dissolve in or react with water. Suggest the approximate pH of the solution formed in **each** case. Explain, with the aid of an equation, the difference between the two values. **[5]**

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