Cambridge University Press

Chemistry

For the IB Diploma

Brighter Thinking
Better Learning

Chapter 10

The periodic table

The periodic table showing the distribution of metals, metalloids and non-metals

period number	group number: 1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 La *	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 Ac **	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

*	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
**	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Figure 10.1: The periodic table showing the distribution of metals, metalloids and non-metals. Nothing much is known about the later elements – they are all produced artificially and only a few atoms have been produced that exist for fractions of a second, but their properties can be predicted from the properties of other elements around them.

Different blocks in the periodic table

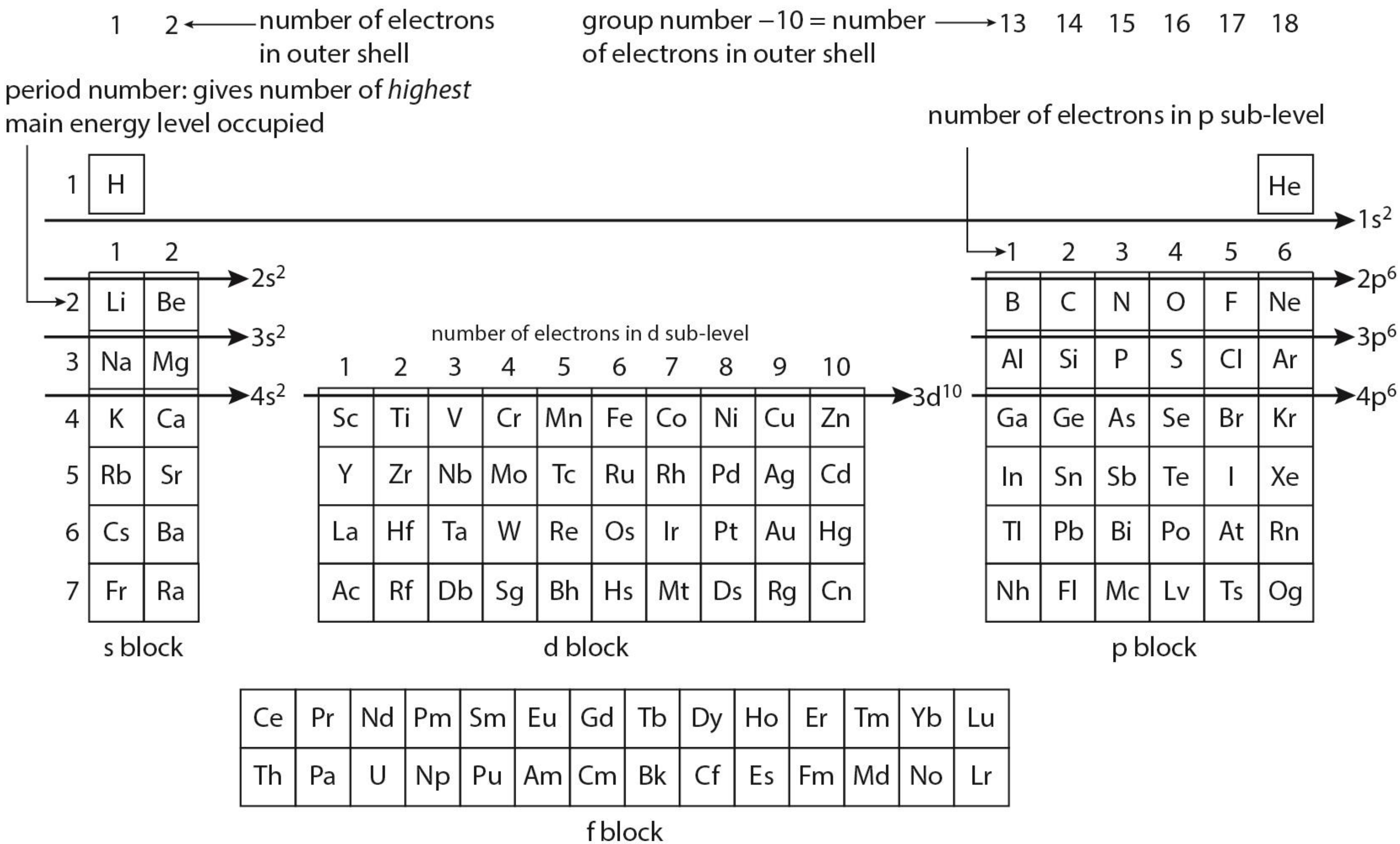

Figure 10.2: Division of the periodic table into blocks.

The atomic radius

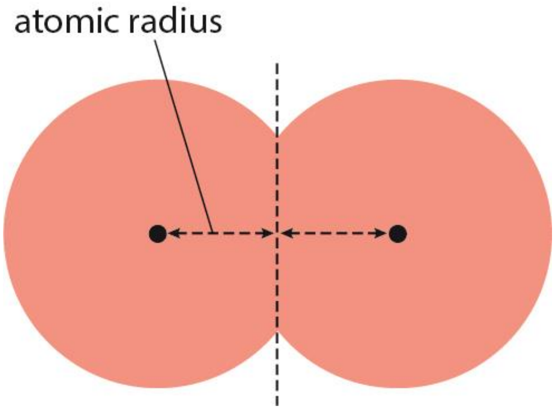

Figure 10.3: Definition of the atomic radius.

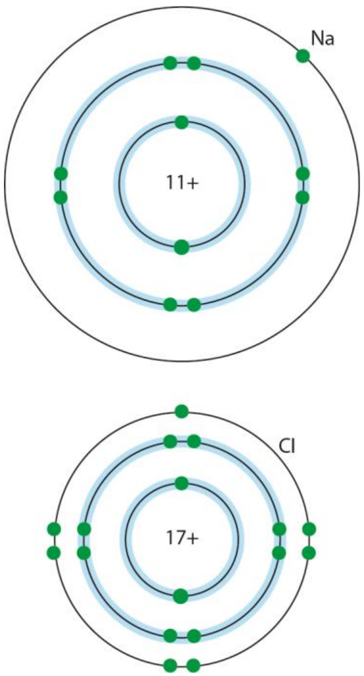

Figure 10.4: Atomic radius decreases across a period. Inner shells, which shield the outer electrons, are highlighted in blue.

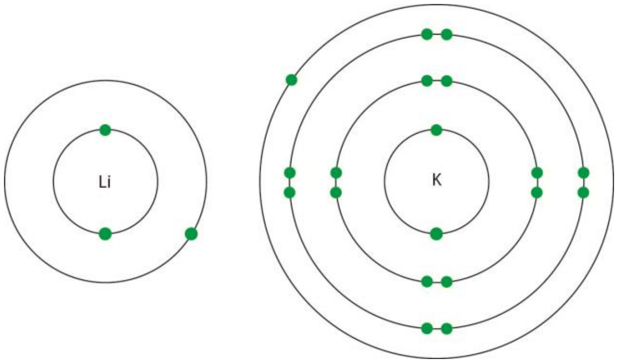

Figure 10.5: Atomic radius increases down a group.

Ionic radii for metals and halogens

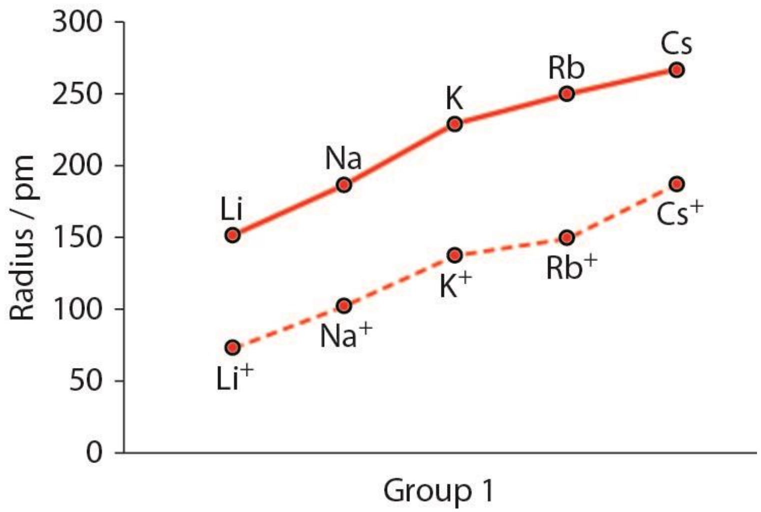

Figure 10.6: Atomic and ionic radii for the alkali metals.

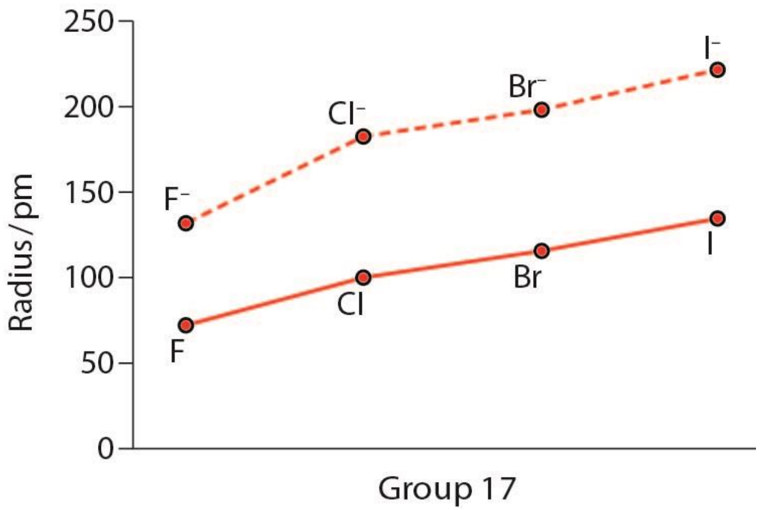

Figure 10.7: A comparison of size between halogen atoms and their ions.

Electron affinity

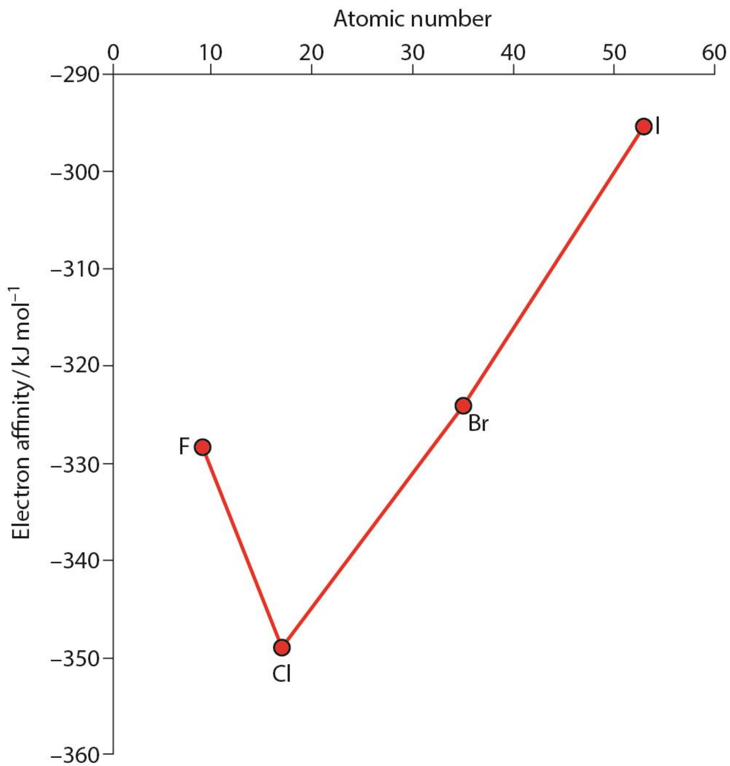

Figure 10.8: Electron affinity values of Group 17 elements.

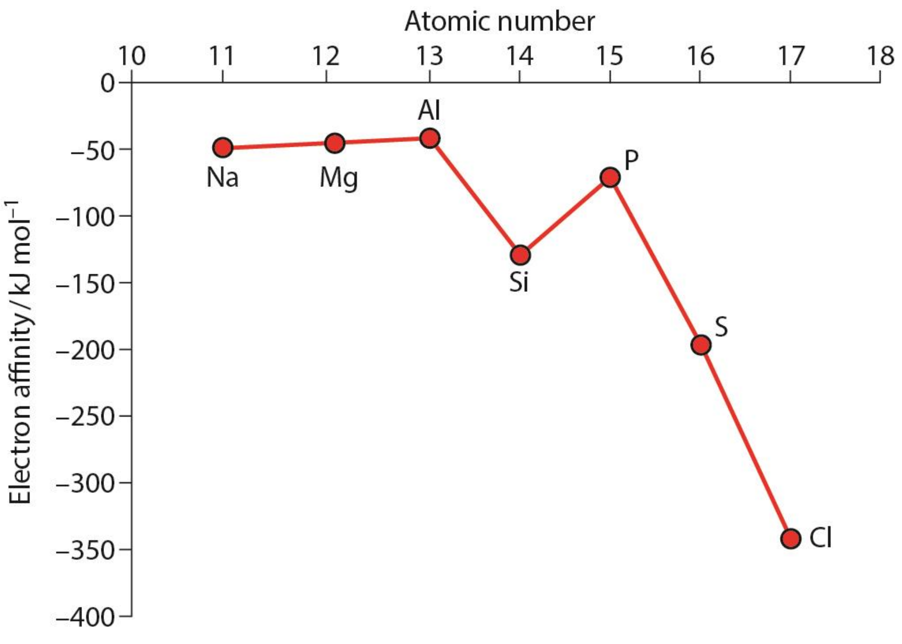

Figure 10.9: Electron affinity values across Period 3.

Electronegativity

Electronegativity: a measure of the attraction of an atom in a molecule for the electron pair in the covalent bond of which it is a part. A more electronegative atom attracts electrons more strongly.

H 2.2							
Li 1.0	Be 1.6		B 2.0	C 2.6	N 3.0	O 3.4	F 4.0
Na 0.9	Mg 1.3		Al 1.6	Si 1.9	P 2.2	S 2.6	Cl 3.2
K 0.8	Ca 1.0		Ga 1.8	Ge 2.0	As 2.2	Se 2.6	Br 3.0
Rb 0.8	Sr 1.0		In 1.8	Sn 2.0	Sb 2.0	Te 2.1	I 2.7

Figure 10.10: Electronegativity values of some elements.

Electronegativity trend

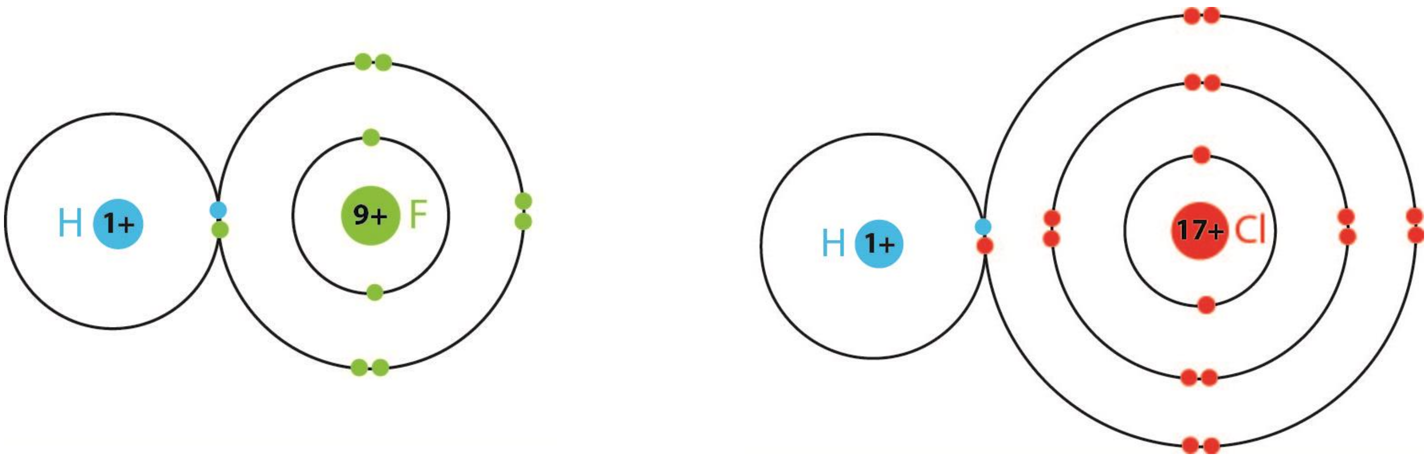

Figure 10.11: Electronegativity decreases down a group.

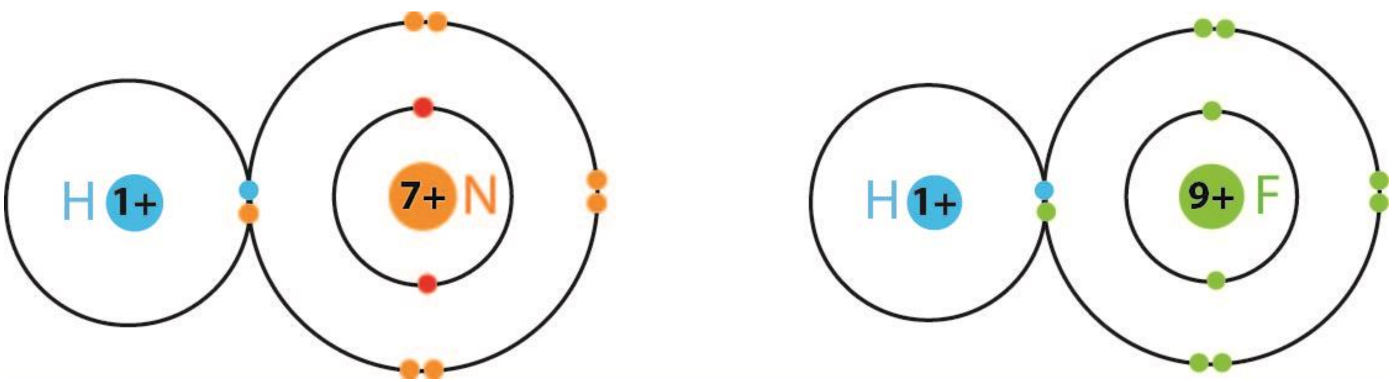

Figure 10.12: Electronegativity increases across a period.

Reactions of the elements in Group 1

The reactions become more vigorous as the group is descended because the ionisation energy decreases as the size of the atom increases.

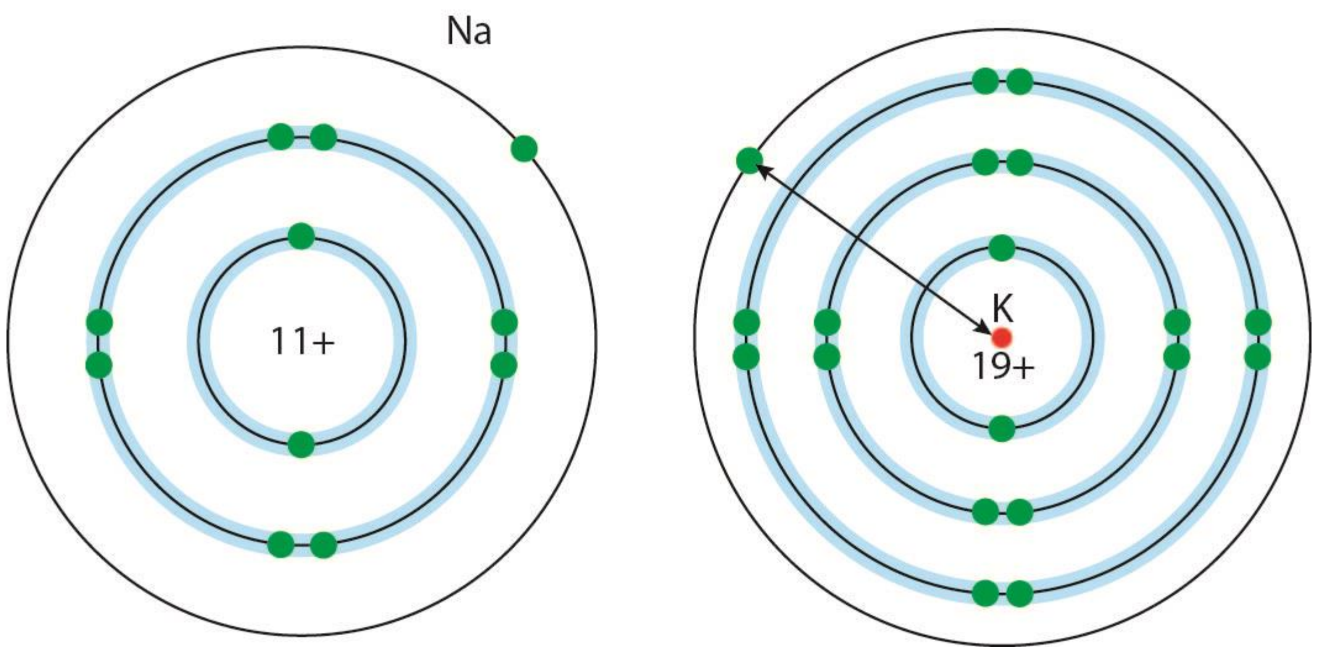

Figure 10.13: Electron configurations of sodium and potassium atoms.

Group 17 displacement reactions

Record the observations of the reactions between halogens and halide compounds.

	KCl(aq)	KBr(aq)	KI(aq)
Cl_2(aq)			
Br_2(aq)			
I_2(aq)			

Oxides of Period 3 elements

Complete the table below.

<table>
<tr><th></th><th>Sodium</th><th>Magnesium</th><th>Aluminium</th><th>Silicon</th><th>Phosphorus</th><th>Sulphur</th><th>Chlorine</th></tr>
<tr><td rowspan="2">Name of Oxide</td><td rowspan="2">Sodium Oxide</td><td rowspan="2">Magnesium oxide</td><td rowspan="2">Aluminium oxide</td><td rowspan="2">Silicon(IV) oxide</td><td>Phosphorus(III) oxide</td><td>Sulphur(IV) oxide</td><td>Chlorine(I) oxide</td></tr>
<tr><td>Phosphorus(V) oxide</td><td>Sulphur(VI) oxide</td><td>Chlorine(VII) oxide</td></tr>
<tr><td rowspan="2">Formula of oxide</td><td rowspan="2"></td><td rowspan="2"></td><td rowspan="2"></td><td rowspan="2"></td><td></td><td></td><td></td></tr>
<tr><td></td><td></td><td></td></tr>
<tr><td rowspan="2">Physical State at 25 °C</td><td colspan="4" rowspan="2">Solid</td><td>Liquid</td><td>Gas</td><td>Gas</td></tr>
<tr><td>Solid</td><td>Liquid</td><td>Liquid</td></tr>
<tr><td>Bonding in oxide</td><td colspan="3">Ionic
Covalent</td><td colspan="4"></td></tr>
<tr><td>Structure</td><td colspan="4">Giant</td><td colspan="3">Simple molecular</td></tr>
<tr><td rowspan="2">Species present in liquid state</td><td rowspan="2"></td><td rowspan="2"></td><td rowspan="2"></td><td rowspan="2"></td><td></td><td></td><td></td></tr>
<tr><td></td><td></td><td></td></tr>
<tr><td>Electrical conductivity when molten</td><td colspan="3"></td><td colspan="4"></td></tr>
</table>

Acid deposition

Problems associated with acid deposition.

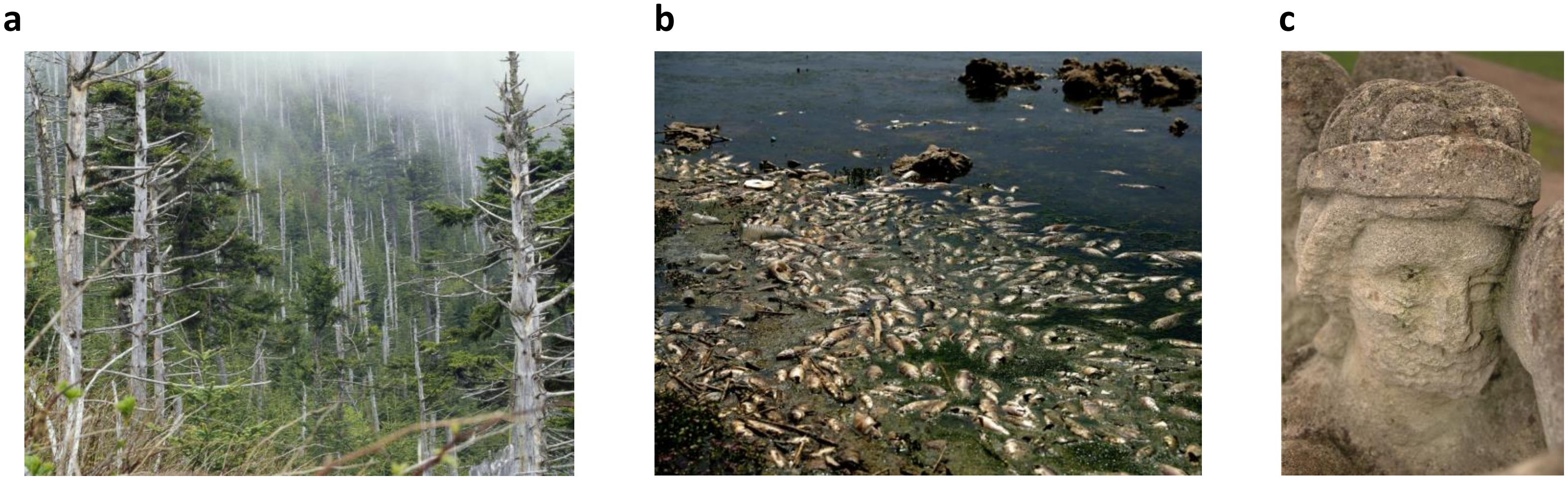

Figure 10.14: Acid rain can **a** kill trees and **b** fish in lakes, and **c** react with limestone buildings to cause corrosion.

Oxidation state

Complete the table with the oxidation states of the elements present in the compound/ions.

SO_2			Cl_2		
SO_3			PCl_3		
H_2O			NH_3		
BaF_2			NF_3		
OF_2			SO_4^{2-}		
Cl_2O_7			ClO_3^-		
NO_3^-			PO_4^{3-}		
$Cr_2O_7^{2-}$			CrO_4^{2-}		
NaH			H_2O_2		

A purely formal concept that regards all compounds as ionic and assigns charges to the elements accordingly.

Oxidation states of compounds

Name the following compounds using oxidation states:

MnO_2	
$CuSO_4$	
$CrCl_3$	
Cr_2O_3	
$Na_2Cr_2O_7$	
$KMnO_4$	
PCl_3	
PCl_5	

A comparison of the variation of first ionisation energy and atomic radius across Period 3 with that across the transition element series

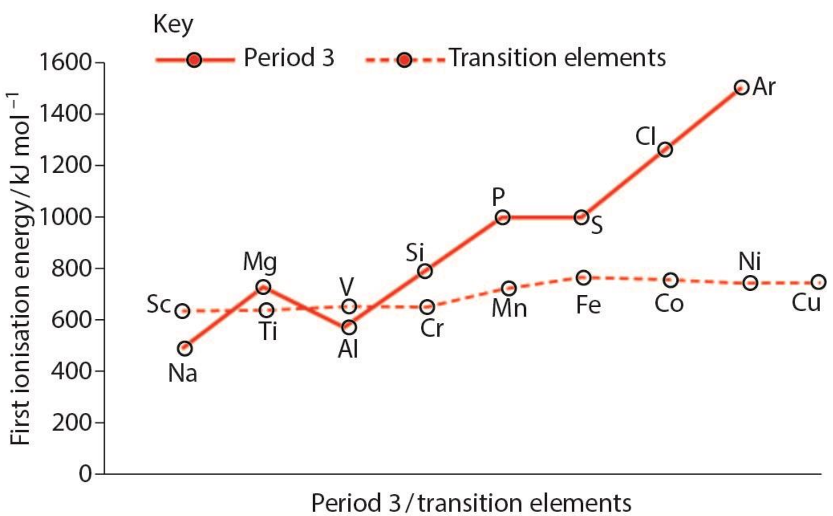

Figure 10.15: A comparison of the variation of first ionisation energy across Period 3 with that across the transition element series.

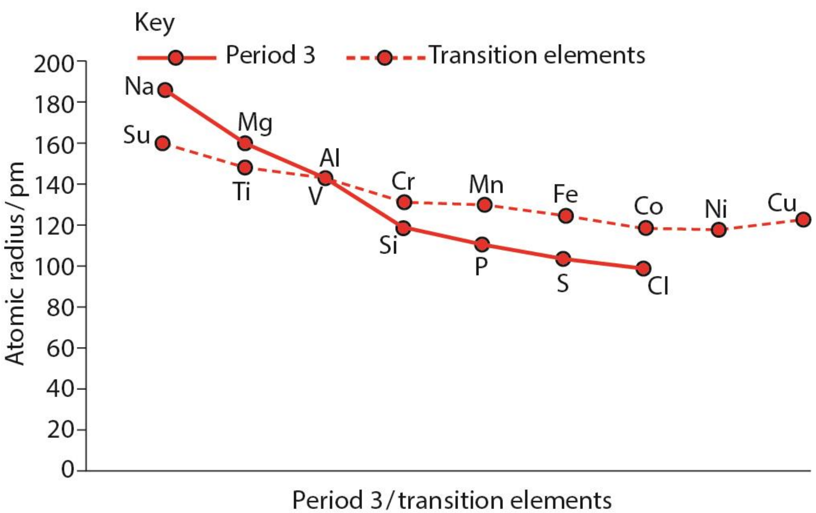

Figure 10.16: A comparison of the variation of atomic radius across Period 3 with that across the transition element series.

Oxidation states of transition elements in compounds

Not all oxidation states are common.

				Mn				
			Cr	+7	Fe			
		V	+6	+6	+6			
	Ti	+5	+5	+5	+5	Co	Ni	Cu
Sc	+4	+4	+4	+4	+4	+4	+4	+4
+3	+3	+3	+3	+3	+3	+3	+3	+3
	+2	+2	+2	+2	+2	+2	+2	+2
	+1	+1	+1	+1	+1	+1	+1	+1
0	0	0	0	0	+0	0	0	0

Figure 10.17: Oxidation states of transition element compounds.

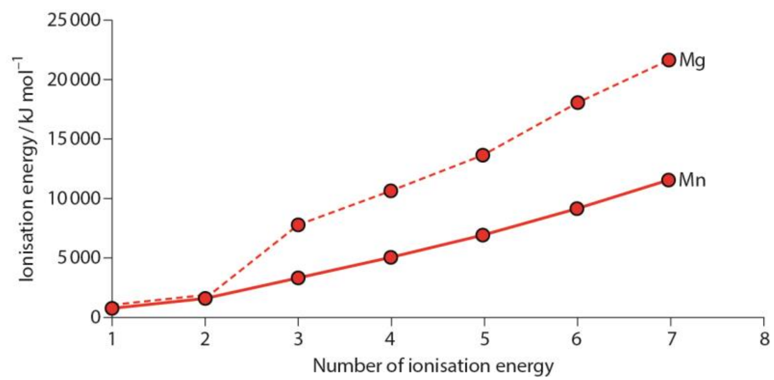

Figure 10.18: Comparison of successive ionisation energies of magnesium and manganese.

Why do transition elements have variable oxidation state?
The 4s and 3d sub-shells are close in energy, and there are no big jumps in the successive ionisation energies when the 4s and 3d electrons are removed.

The oxidation state of a transition element in a complex ion

$[Fe(H_2O)_6]^{2+}$

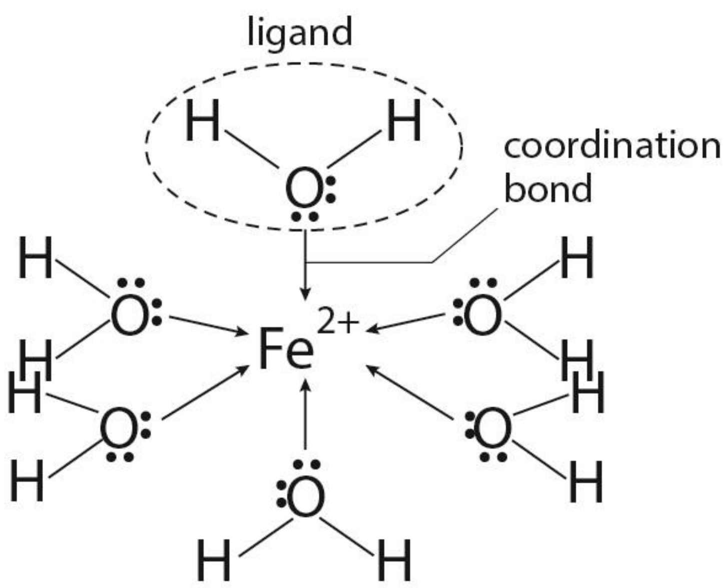

Figure 10.19: A complex ion is formed when ligands bond to a transition element ion. The ligands donate lone pairs into vacant orbitals (3d, 4s or 4p) on the transition element ion.

Neutral ligands	1^- ligands
H_2O	Cl^-
NH_3	CN^-

Complete this table

	Oxidation state of transition element
$[Fe(H_2O)_6]^{2+}$	
$[Ni(CN)_4]^{2-}$	
$[Co(NH_3)_2Cl_4]^-$	