Oxford Resources for IB

Chemistry – 2023 Edition

Answers

Structure 2.1 – The ionic model

Activity (page 97)

S: number of protons = 16, number of electrons = 16, so sum of charges = 0, $1s^2\,2s^2\,2p^6\,3s^2\,3p^4$;

S^{2-}: number of protons = 16, number of electrons = 18, so sum of charges = –2, $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$

Practice questions (page 100)

1. **a.** 1+
   b. 2+
   c. 3+
   d. 1–
   e. 3–
   f. 2–
   g. 2+
2. **d.** fluoride
   e. nitride
   f. selenide
3.

Name	Symbol	Number of protons	Number of electrons	Electron configuration	Charge
beryllium	Be	4	4	$1s^2\,2s^2$	0
potassium	K^+	19	18	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$	1+
oxide	O^{2-}	8	10	$1s^2\,2s^2\,2p^6$	2–
phosphide	P^{3-}	15	18	$1s^2\,2s^2\,2p^6\,3s^2\,3p^6$	3–
hydrogen	H^+	1	0	N/A	1+

4. **a.** Any three species with electron configuration $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$, e.g. Ar, Cl^-, S^{2-}, P^{3-}, K^+, Ca^{2+}, Sc^{3+}
   b. Any three species with electron configuration $1s^2$, e.g. He, H^-, Li^+, Be^{2+}
5. Noble gas atoms do not gain or lose electrons easily and hence do not form ions. They have very high ionization energies. They have full (or “closed”) outer electron shells.

Practice questions (page 101)

6. **a.** $[Ar]\,3d^5$
   b. $[Ar]\,3d^2$
   c. $[Ar]\,3d^{10}$
   d. $[Ar]\,3d^9$

7. a. $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}$
   b. Transition elements form at least one ion that has an incomplete d subshell. This is not the case for zinc. The electron configuration of zinc is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}\,4s^2$. Zinc atoms readily form Zn^{2+} ions by losing the two 4s electrons, but the 3d sub-shell remains complete with 10 electrons. The electron configuration of a Zn^{2+} ion is $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}$.

8. a. ${}^{55}_{25}Mn^{2+}$
   b. Cr^+

Data-based question (page 103)

The difference in electronegativity between sodium and fluorine is much greater than 1.8. therefore, sodium fluoride is ionic. Sodium chloride is ionic for similar reasons.

Activity (page 104)

a. i. Li and F are in groups 1 and 17, respectively. On the periodic table, they are horizontally far apart from each other and are therefore likely to bond ionically.
   ii. χ(Li) = 1.0 and χ(F) = 4.0; $\Delta\chi = 3.0$; $\Delta\chi > 1.8$; therefore Li and F bond ionically

b. i. Rb and Ga are in groups 1 and 13, respectively. On the periodic table, they are horizontally not very far apart from each other and are therefore unlikely to bond ionically.
   ii. χ(Rb) = 0.8 and χ(Ga) = 1.8; $\Delta\chi = 1.0$; $\Delta\chi < 1.8$; therefore Rb and Ga do not bond ionically

c. i. Ca and I are in groups 2 and 17, respectively. On the periodic table, they are horizontally far apart from each other, suggesting they bond ionically. However, electronegativity decreases down groups, and because they are both fairly low in their groups, the difference in electronegativity between them is likely to be close to 1.8. This suggests a bond with both ionic and covalent characteristics.
   ii. χ(Ca) = 1.0 and χ(I) = 2.7; $\Delta\chi = 1.7$; $\Delta\chi < 1.8$, therefore Ca and I do not bond ionically but, because $\Delta\chi$ is close to 1.8, the bond between them will have a considerable percentage of ionic character.

d. i. As and S are in groups 15 and 16, respectively. On the periodic table, they are horizontally close to each other and are therefore unlikely to bond ionically.
   ii. χ(As) = 2.2 and χ(S) = 2.6; $\Delta\chi = 0.4$; $\Delta\chi < 1.8$, therefore As and S do not bond ionically

e. i. P and Cl are in groups 15 and 17, respectively. On the periodic table, they are horizontally close to each other and are therefore unlikely to bond ionically.
   ii. χ(P) = 2.2 and χ(Cl) = 3.2; $\Delta\chi = 1.0$; $\Delta\chi < 1.8$, therefore P and Cl do not bond ionically

f. i. Ag and Br are in groups 11 and 17, respectively. On the periodic table, the horizontal distance between them is intermediate, suggesting they might bond ionically to an extent. However, electronegativity decreases down groups, and because they are both fairly low in their groups, the difference in electronegativity between them is likely to be close to 1.8. This all suggests a bond with both ionic and covalent characteristics.
   ii. χ(Ag) = 1.9 and χ(Br) = 3.0; $\Delta\chi = 1.1$; $\Delta\chi < 1.8$, therefore Ag and Br do not bond ionically on the basis of the $\Delta\chi$. However, as we will see in *Structure 2.4*, the mean electronegativity of the two elements also plays a part in determining the type of bond between them. Silver bromide in fact exhibits both ionic and covalent character.

Practice questions (page 105)

9. a. rubidium fluoride
   b. aluminium sulfide
   c. aluminium nitride
   d. strontium hydroxide
   e. barium carbonate
   f. ammonium hydrogencarbonate

Practice questions (page 107)

10. a. MgO

b. $SrCl_2$

c. Na_2S

d. Li_3N

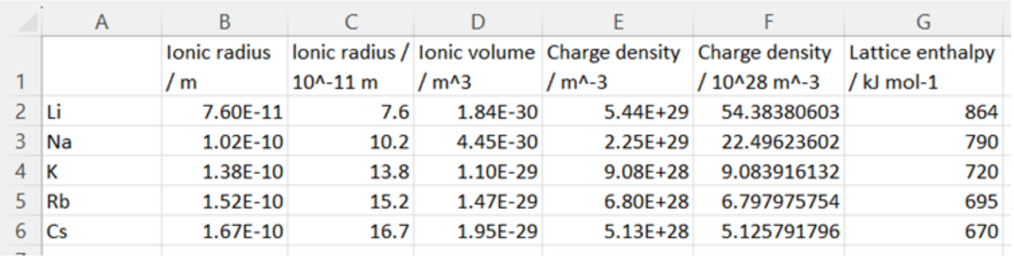
e. $LiNO_3$

f. $Ba(HCO_3)_2$

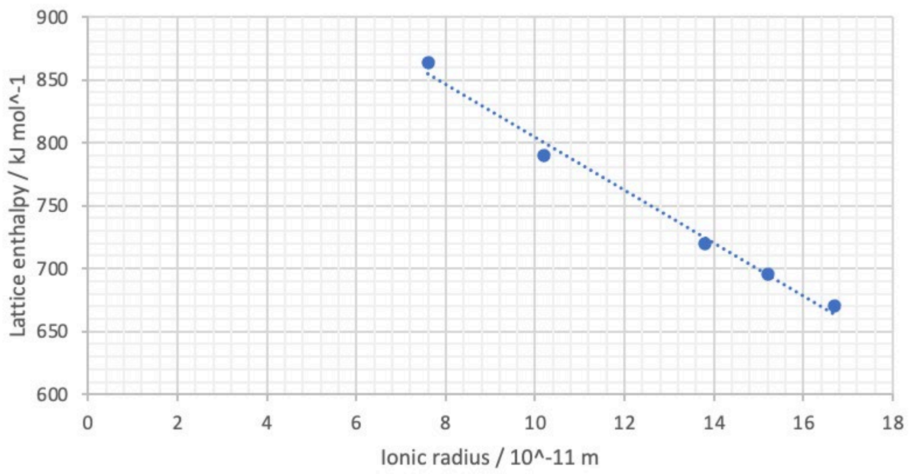
g. $(NH_4)_3PO_3$

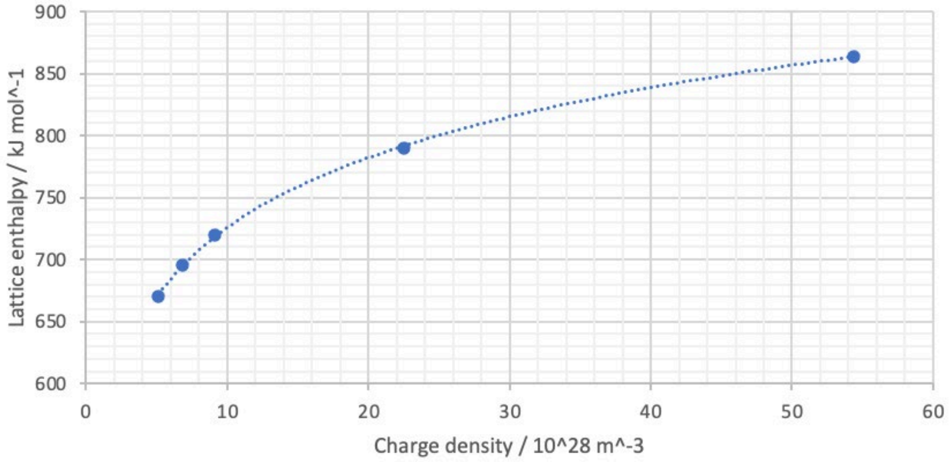
Skills questions (page 110)

5, 6, 7. For example:



	A	B	C	D	E	F	G
1		Ionic radius / m	Ionic radius / 10^-11 m	Ionic volume / m^3	Charge density / m^-3	Charge density / 10^28 m^-3	Lattice enthalpy / kJ mol-1
2	Li	7.60E-11	7.6	1.84E-30	5.44E+29	54.38380603	864
3	Na	1.02E-10	10.2	4.45E-30	2.25E+29	22.49623602	790
4	K	1.38E-10	13.8	1.10E-29	9.08E+28	9.083916132	720
5	Rb	1.52E-10	15.2	1.47E-29	6.80E+28	6.797975754	695
6	Cs	1.67E-10	16.7	1.95E-29	5.13E+28	5.125791796	670





8. Answers will vary. For example the first graph shows a linear negative relationship, suggesting that an increased ionic radius results in a lower lattice enthalpy, thus, the ionic bond is weaker by roughly 60 kJ mol^{-1}. The second graph shows a non-linear positive relationship, suggesting that an ionic charge distributed over a smaller volume results in a higher lattice enthalpy. The relationship is linear for the first three alkali metals and presents a less drastic impact as the charge density increases for Rb and Cs.

9. The first graph is linear. The second is not. Both show that larger ions result in a lower lattice enthalpy, but the second takes charge per unit volume into account. Because the volume is proportional the cube of the radius, larger radii exhibit significantly lower lattice enthalpy.

10. Answers will vary but you should consider the linearity, domain and range of the data in the graphs you sketched.

11. Answers will vary.

Practice questions (page 110)

1. $KBr(s) \rightarrow K^+(g) + Br^-(g)$

   $CaO(s) \rightarrow Ca^{2+}(g) + O^{2-}(g)$

   $MgCl_2(s) \rightarrow Mg^{2+}(g) + 2Cl^-(g)$
2. KF has a lower lattice enthalpy value. Both KF and K_2O contain K^+ ions. The difference must be due to the anions, F^- and O^{2-}. The charge of O^{2-} ions is greater and therefore the electrostatic attraction between oxide ions and potassium ions is greater than that between fluoride ions and potassium ions. The effect of charge is more significant than the effect of the fluoride ion having a slightly smaller radius than the oxide ion.
3. MgO has the highest lattice enthalpy out of the compounds listed. The electrostatic attraction between oppositely charged ions is stronger for two reasons: smaller ionic radius and greater charge. The ionic radii of Mg^{2+} and O^{2-} are smaller, compared to Na^+ and Cl^-, respectively. The charges of the magnesium and oxide ions are also greater than the sodium and chloride ions.
4. The lattice enthalpy of the group 1 chlorides decreases down the group from LiCl to CsCl. This is because ionic radii increase down the group due to the presence of an additional energy level. Larger ionic radii are associated with lower lattice enthalpy values because of the decreasing strength of the electrostatic attractions between oppositely charged ions that results.

Activity (page 112)

Solid: diagram showing a regular lattice of positive and negative ions that cannot change position, even when a potential difference is applied.

Molten: a diagram showing a random arrangement of positive and negative ions that can move past each other when a potential difference is applied.

Aqueous: a diagram showing a random arrangement of positive and negative ions, and water molecules, all of which can move past each other when a potential difference is applied.

Skills questions (page 114)

A possible method is:

1. Dip the copper wire into samples of each of the three solutions. Only silver nitrate will react in a displacement reaction.
2. Once identified, silver nitrate can be added to samples of the other solutions. The one that does not form a precipitate is calcium nitrate.
3. Add acid to the remaining two. Potassium carbonate will effervesce when acid is added.
   The remaining, unidentified solution is sodium chloride.

End of topic questions (pages 115–116)

1. Answers will depend on the student, but a possible answer might be: The extent to which a bond will be ionic depends on the difference in the electronegativity of the substances in the bond. The greater the difference in electronegativity, the greater the ionic character.

   Ionic bonds are the electrostatic attraction between cations and anions. It is therefore natural that the presence and forces of attraction between these ions should dictate many of the properties of ionic compounds, such as their lattice structure and enthalpy. Ionic structures are non-volatile. They are poor electrical conductors when solid due to the absence of mobile charged particles. When aqueous or molten, however, ionic compounds conduct electricity because their component ions are able to move.

It is often said that ionic bonds are formed when a metal bonds to a non-metal element, but this is not entirely correct. The root of this generalization comes from the fact that metals and non-metals are typically found on opposite ends of the periodic table, hence they have a large difference in electronegativity.

2. Ionic substances are always compounds because ionic bonds require there to be a significant difference in the electronegativity of the elements involved in the bond. Ionic substances cannot be elements. When two atoms of the same element bond together, the difference in electronegativity is zero. Therefore, when two atoms of the same element bond, the bond formed will either be metallic or covalent, but never ionic.

3. C

The atoms of the group 17 elements have 7 outer electrons and therefore generally gain one electron forming 1– anions.

4. A

If the ion contains 20 protons its atom must contain 20 electrons (it is calcium). The ion's 2+ charge indicates that it has lost two electrons. The total number of electrons is 18. The other answers contain the wrong number of electrons.

5. A

A is the correct answer. B is incorrect because ionic compounds are not made up of molecules. C is incorrect because electrons are not delocalized in ionic compounds. D is incorrect because what holds oppositely charged ions together in an ionic lattice is not the electron transfer that may have led to the formation of the ions, but rather the electrostatic attraction between cations and anions resulting from them having opposite charges.

6. D

Ca is calcium and SO_4 indicates the compound contains a sulfate ion. Note: sulfite is the non-systematic name for SO_3^{2-}, which has systematic name sulfate(IV) due to the +4 oxidation state of the sulfur atom.

7. B

Sodium forms Na^+ ions. Since nitrate ions have a –1 charge, a compound containing Na^+ ions and NO_3^- ions in a 1:1 ratio will have zero overall charge. Note: nitrite is the non-systematic name for NO_2^-, which has systematic name nitrate(III) due to the +3 oxidation state of the nitrogen atom.

8. B

Lattice enthalpy increases with increasing charge and decreasing ionic radius. Calcium ions, Ca^{2+}, have a greater charge than potassium ions, K^+, and therefore the calcium compounds are likely to have higher lattice enthalpies. Oxide ions have a smaller radius than sulfide ions because they have one fewer energy level. Out of the compounds listed, CaO is likely to have the strongest electrostatic attractions between its ions and, thus, the greatest lattice enthalpy.

9. D

Lattice enthalpy increases with increasing charge and decreasing ionic radius. These conditions lead to a greater charge density on the ions, resulting in a greater electrostatic attraction between oppositely charged ions. The relationships stated in A and C are the wrong way around. Answer B gives an incorrect definition of lattice enthalpy.

10. A

Lattice enthalpies represent the process of overcoming electrostatic attractions between cations and anions in a solid lattice, producing gaseous ions.

Answer B is incorrect because it shows the solid lattice breaking down into component standard-state elements. Answer C gives the potassium ion as diatomic, which is incorrect. Answer D is incorrect because its products are not ions.

11. D

Lattice enthalpy increases with decreasing ionic radius. Ionic radius increases down groups. In the case of the halides, ionic radius increases from fluoride to iodide. Out of the lithium halides, lithium fluoride will have the greatest lattice enthalpy and should therefore be listed last (because the question asks for them in order of increasing lattice enthalpy). The answer is therefore C. Answer D has bromide and chloride the wrong way round.

12. C

The properties of a typical ionic compound are high melting point and electrical conductivity when molten or dissolved. Ionic compounds are often soluble in water, but not always, and therefore aqueous solubility is not a reliable indicator. When solid, ionic compounds do not conduct electricity. Answer C is the only one that contains a substance that does not conduct electricity when solid, but does when molten. It also has a high melting point.

13. a. CaF_2

b. ionic bonding;

electrostatic attraction between (Ca^{2+}) cations and (F^-) anions / oppositely charged ions

(arranged in a) lattice structure;

c. Ionic compounds do not conduct electricity when solid because ions cannot move / are held in place;

When the substance is molten or dissolved, the ions are mobile (when a potential difference is applied)

d. i. $CaF_2(s) \rightarrow Ca^{2+}(g) + 2F^-(g)$

M1: correct formulas and balancing; *M2: correct state symbols*

ii. (Lattice enthalpies represent) the energy required to overcome the electrostatic attractions between oppositely charged ions in an ionic solid;

This process absorbs energy and is therefore endothermic;

iii. The compounds contain different anions: one contains fluoride ions, while the other contains oxide ions. The 1– charge of fluoride ions is lower in magnitude than the 2– charge of oxide anions;

Therefore the electrostatic attraction between calcium ions and oxide ions in CaO is greater than that between calcium ions and fluoride ions in CaF_2. Overcoming the greater electrostatic attraction requires a greater energy and thus CaO has a greater lattice enthalpy.

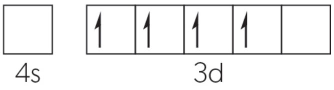
14. a. K^+: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6$

b. There is a one electron difference between the two species. The mass of an electron (1/2000 that of a proton) is negligible in this context. The masses of a potassium ion and a potassium atom will therefore be practically identical (assuming that their mass numbers are equal, of course).

c. The dichromate(VI) ion has a 2– charge: $Cr_2O_7^{2-}$

d. i. Cr: $[Ar]\,4s^1\,3d^5$

ii.



iii. Cr^{3+}: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^3$

15. a. $2Li(s) + F_2(g) \rightarrow 2LiF(s)$

b. 1+

c. Oxidized: lithium (because it loses electrons when it becomes a positively charged lithium ion)

Reduced: fluorine (because it gains electrons when it becomes a negatively charged fluoride ion)

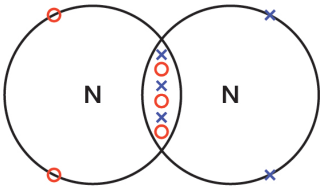
d. A diagram showing a 3D continuous network of alternating lithium ions and fluoride ions;

where the radius of the lithium ions is smaller than that of the fluoride ions

Structure 2.2 – The covalent model

Activity (page 119)

A diagram such as this is needed:

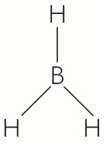


Practice questions (page 119)

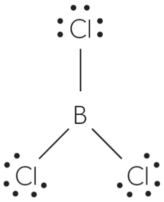
1. **a.** carbon and oxygen: $\Delta\chi = 3.4 - 2.6 = 0.8$, covalent
   b. sodium and oxygen: $\Delta\chi = 3.4 - 0.9 = 2.5$, ionic
   c. carbon and hydrogen: $\Delta\chi = 2.6 - 2.2 = 0.4$, covalent
   d. chlorine and oxygen: $\Delta\chi = 3.4 - 3.2 = 0.2$, covalent
   e. iodine and iodine: $\Delta\chi = 2.7 - 2.7 = 0$, covalent
   f. aluminium and fluorine: $\Delta\chi = 4.0 - 1.6 = 2.4$, ionic

Practice questions (page 123)

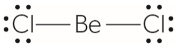
2. **a.**



b.



c.



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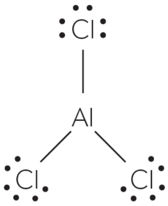
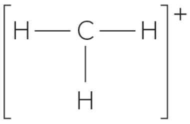
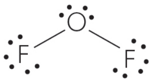


e.



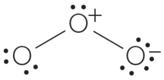
3. **a.**

b.



c.

d.



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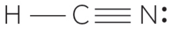
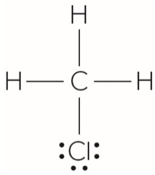
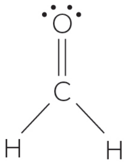


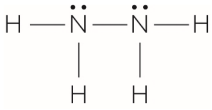
f.



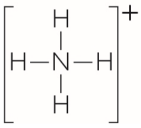
g.



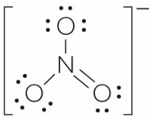
h.



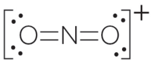
4. **a.**



b.



c.



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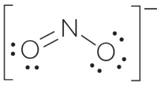


e.

Data-based question (page 125)

Answers will vary, depending on your method for estimating the bond lengths and bond enthalpies.

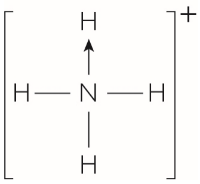
Guidance: the reasoning and assumptions used to reach the estimates should be expressed clearly. Resulting values should be compared to the corresponding data booklet values.

For example,

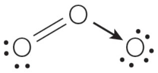
- Single carbon-nitrogen bond strength: 367 kJ mol^{-1} (Which is 2.42 times lower than the given C–N bond strength. This is the ratio between the given carbon-carbon triple and single bond data.) The data booklet value is 286 kJ mol^{-1} giving a 28.3% error.
- Double carbon-nitrogen bond strength: 651 kJ mol^{-1} (1.37 times lower than the given C–N bond strength.) The data booklet value is 615 kJ mol^{-1}, so this percentage error is 5.85%.

Practice questions (page 126)

5. a.

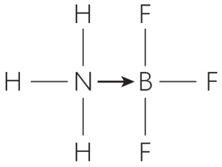


b.



c.

d.



Skills questions (page 131)

A wide range of answers possible, depending on the database(s) and molecules chosen.

Practice questions (pages 131)

6. a. EDG: tetrahedral; MG: tetrahedral

b. EDG: trigonal planar; MG: bent or V-shaped

c. EDG: linear; MG: linear

d. EDG: tetrahedral; MG: bent or V-shaped

e. EDG: tetrahedral; MG: trigonal pyramidal

7. a. two bonding domains, zero non-bonding domains

b. three bonding domains, one non-bonding domain

c. four bonding domains, zero non-bonding domains

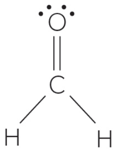
d. three bonding domains, zero non-bonding domains

e. two bonding domains, either one or two non-bonding domains

8.
   a. trigonal pyramidal
   b. tetrahedral
   c. trigonal planar
   d. linear
   e. linear
   f. trigonal planar
   g. bent or V-shaped
   h. bent or V-shaped

9.
   a. <109.5°, e.g. 107°
   b. approximately 109.5° (a slight difference will be caused by different sizes of H and Cl atoms)
   c. 120°
   d. 180°
   e. 180°
   f. 120°
   g. <109.5°, e.g. 104.5°
   h. <120°, e.g. 117°

10.
   a.



   b. trigonal planar
   c. ∠OCH = slightly greater than 120°, e.g. 122° and ∠HCH is slightly smaller than 120°, e.g. 116°. The double bond contains four electrons and therefore repels the other two domains strongly, increasing the ∠OCH bond angles and, subsequently, reducing the ∠HCH bond angle.

Practice questions (page 132)

11. From least polar to most polar: C–C < C–H < C–O < C–F

12. Br_2 is non-polar because the difference in electronegativity between its two atoms is zero; this is because the two atoms are identical. HBr is polar because H and Br have different electronegativities (2.2 and 3.0, respectively). They form a covalent bond with a partial negative charge on the bromine and a partial positive charge on the hydrogen.

Practice questions (page 134)

13.
   a. non-polar
   b. polar
   c. polar
   d. (almost) non-polar

14. The C=O bonds are polar but the molecule is linear and therefore the two bond dipoles cancel out. There is no net dipole.

15. Phosphine, PH_3, is slightly polar. P and H have the same electronegativity values, hence the bonds are non-polar. However, the molecule is trigonal pyramidal due to the presence of a lone pair of electrons. PH_3 is therefore polar because the electrons in the molecule are distributed unevenly.

Practice questions (page 137)

16. Graphite and graphene contain delocalized electrons. These electrons can move through the structure when a potential difference is applied to it. In diamond, the electrons are localized in the covalent bonds between carbon atoms. Diamond is an electrical insulator due to the absence of mobile charged particles.

17. Diamond and silicon dioxide are examples of substances with covalent network structures. Their atoms are bonded covalently in an extended 3D lattice. Melting and boiling covalent networks such as diamond and silicon dioxide requires breaking many covalent bonds. Covalent bonds are very strong and require a large amount of energy to break, therefore the melting and boiling points of these substances are high.

Activity (page 138)

	Diamond	Graphite	Graphene	C_{60} fullerene	Silicon dioxide
Element(s)	carbon	carbon	carbon	carbon	silicon and oxygen
Arrangement of atoms	3D lattice of covalently bonded tetrahedral carbon atoms	layers of covalently bonded trigonal planar atoms in a hexagonal honeycomb arrangement	a single layer of covalently bonded trigonal planar atoms in a hexagonal honeycomb arrangement	spherical molecular structure composed of 60 covalently bonded and trigonal planar carbon atoms arranged in an array of 20 hexagons and 12 pentagons	3D lattice of tetrahedral silicon atoms bonded to oxygen atoms in a ratio of 1 to 2
Electrical conductivity	poor	high	high	poor	poor

Data-based questions (page 142)

1. The boiling points of the first four alkanes increase with increasing molecular size due to the increasing number of electrons. The larger electron clouds are more easily polarizable, strengthening the London dispersion forces between molecules. Stronger London dispersion forces require more energy to be overcome and hence the boiling points increase.
2. Pentane, 2-methylbutane and 2,2-dimethylpropane are isomers with molecular formula C_5H_{12} and identical molecular mass, but different branching and hence different shapes. Pentane is a straight-chain alkane. 2-methylbutane has one alkyl branch and 2,2-dimethylpropane has two alkyl branches. Because of its intermediate branching, the boiling point of 2-methylbutane is likely to be between the other two boiling points. One way to roughly estimate this is finding the midpoint between 9.5 °C and 36.1 °C, giving 22.8 °C. (The reported boiling point of 2-methylbutane is 30 °C. The fact that the estimated boiling point and the actual boiling point are different suggests a non-linear relationship between branching and boiling point.)

Activity (page 142)

Both are types of intermolecular force and involve a molecule inducing a dipole on a neighbouring molecule.

London (dispersion) forces only:

- an instantaneous and temporary dipole induces a dipole on a neighbouring molecule
- occur between all molecules
- non-polar molecules experience only this type of intermolecular force and no other

Dipole–induced dipole forces only:

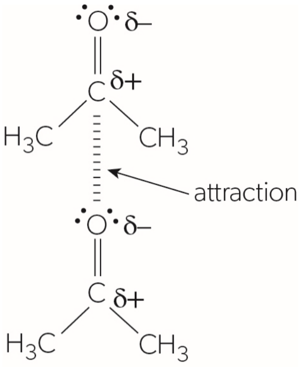
- a permanent dipole induces a dipole on a neighbouring molecule
- occur between a polar molecule and a non-polar molecule

Practice questions (page 143)

18. C. Carbon monoxide, CO, because it is polar

19. The molecular masses of ICl and Br_2 are comparable: 162.35 g mol^{-1} and 159.80 g mol^{-1}. They therefore have similar numbers of electrons and experience LDFs of similar strength between their molecules. However, the presence of a permanent dipole in polar ICl means that, as well as LDFs, dipole–dipole forces also exist between its molecules, adding to the overall intermolecular force strength. This increases the amount of energy needed to overcome the intermolecular forces and therefore its boiling point is higher.

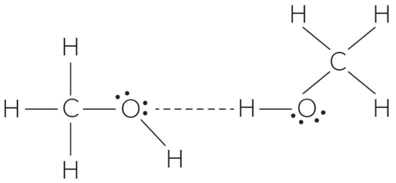
20. Propanone has a permanent dipole, where the oxygen has a partial negative charge and carbon atoms have a partial positive charge. Any arrangement of two propanone molecules in which opposite partial charges are aligned, will be acceptable answers to this question. For example:



Practice questions (page 144)

21. Hydrogen fluoride (A) and methanol (C)

22.



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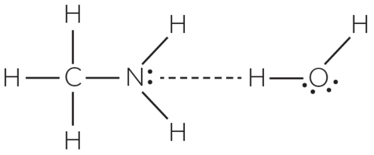
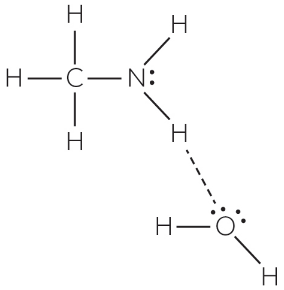


or



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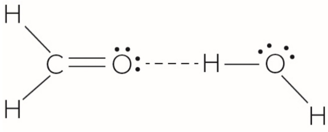


Practice questions (page 149)

25. a. London (dispersion) forces

b. Dipole–dipole forces and London (dispersion) forces

c. Hydrogen bonding, dipole–dipole forces and London (dispersion) forces

d. London (dispersion) forces

e. Dipole–induced dipole forces and London (dispersion) forces

f. Dipole–dipole forces and London (dispersion) forces

26. a. Ammonia can form hydrogen bonds with water because it contains an electronegative atom that is directly bonded to a hydrogen atom

b. Propane is non-polar and therefore cannot form hydrogen bonds

c. Ethanoic acid can form hydrogen bonds with water because it contains an electronegative atom that is directly bonded to a hydrogen atom

Data-based question (page 150)

Answers will vary e.g.

- Boiling point increases with increasing intermolecular force strength
- Molecular dipole moment is low for non-polar substances.
- Molecular dipole moment is not necessarily highest for higher-boiling point substances
- Molecules that can form multiple hydrogen bonds with other like molecules (e.g. methanoic acid) have higher boiling points.

Practice questions (page 151)

27. a. Kr, because it has more electrons

b. NH_3, because it can form hydrogen bonds and is polar

28. We can organize the information in a table as follows. Even though fluoromethane is polar and hence forms dipole–dipole forces between its molecules, non-polar carbon tetrachloride has a higher boiling point because it forms stronger London (dispersion) forces due to the presence of more electrons, as indicated by its significantly larger molar mass in comparison to that of fluoromethane.

Compound	Boiling point / °C	Molar mass / g mol^{-1}	Intermolecular forces
CCl_4	76.7°	153.82	London (dispersion) forces
CH_3F	–78.2	34.03	Dipole–dipole and London (dispersion) forces

Skills questions (page 154)

Answers will vary. Answers could include:

- Step 1: Add distilled water to samples of A, B and C, and stir. White sand will not dissolve, whereas the other two will.
- Step 2: Prepare solutions of the two soluble solids and measure their conductivity, and that of distilled water, with a probe. NaCl(aq) will have a much larger conductivity than the sugar solution, whose conductivity will be only a little larger than that of distilled water.

Many other options are possible. For example, conductivity data can also be obtained by electrolysing the two solutions in step 2 and watching for the formation of bubbles at the two electrodes. Alternatively, sugar will melt if placed in a Bunsen flame, whereas sodium chloride will not.

The tests should not include taste tests.

Practice questions (page 154)

29. Most covalent compounds lack mobile charged particles and therefore do not conduct electricity well. Their electrons are localized in bonds. Graphite and graphene are an electrical conductors because they contain delocalized electrons.

30. The more volatile a substance, the more easily it vaporizes or turns into a gas. Therefore the higher the boiling point, the lower the volatility. High boiling point and low volatility are consequences of strong intermolecular forces that require a significant amount of thermal energy to break.

31. Covalent network substances are not composed of molecules. In order to melt them or boil them, covalent bonds need to be overcome. Covalent bonds are very strong and therefore require a lot of thermal energy to break. Molecular substances contain molecules. Melting and boiling in molecular substances are brought about by the breaking of intermolecular forces, not (intramolecular) covalent bonds. Since intermolecular forces are weak in comparison to covalent bonds, they do not require as much energy to break. Consequently the melting points and boiling points of molecular substances are lower than those of covalent network substances.

32. Ethanol contains a hydrogen directly bonded to an oxygen. It can therefore associate strongly with water molecules through the formation of hydrogen bonds with water molecules, allowing ethanol and water to be miscible.

33. A. ethanoic acid, CH_3COOH, D. propanone, CH_3COCH_3

34. Oil is mainly composed of hydrocarbon chains, which are non-polar. Ionic substances contain ions, which are strongly attracted to one another through electrostatic attraction. Ions and non-polar molecules do not associate strongly and therefore ionic substances do not dissolve in non-polar solvents.

35. Bromine is non-polar and therefore does not associate greatly with water molecules; water molecules are highly polar and are strongly associated with each other through hydrogen bonding. Hexane molecules are held together by weak London (dispersion) forces and can therefore easily associate with a bromine molecule.

Skills questions (page 159)

Answers will vary, depending on the leaf species, stationary phase chosen, composition of the mobile phase, etc.

Practice questions (page 160)

36. 0.50 (note that the question erroneously refers to figure 58 instead of figure 57)

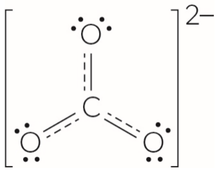
37. Pencil marks do not dissolve in the solvent, allowing one to identify the baseline even when the paper is wet.

38. Mixture A contains the red (bottom small spot) and green (top large spot) pigments, which correspond to individual pigments C and B.

Practice questions (page 162)

39.

40.

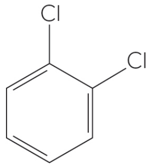


The carbon–oxygen bond order is 1.33 because there are four electron pairs distributed across three domains: 4 ÷ 3 = 1.33

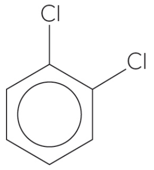
41. The carbon–oxygen bond order is 1.33 in carbonate, 2 in carbon dioxide and 1 in methanol. In order of strength of the carbon–oxygen bond: methanol < carbonate < carbon dioxide. In order of length of the carbon-oxygen bond: carbon dioxide < carbonate < methanol.

Activity (page 163)

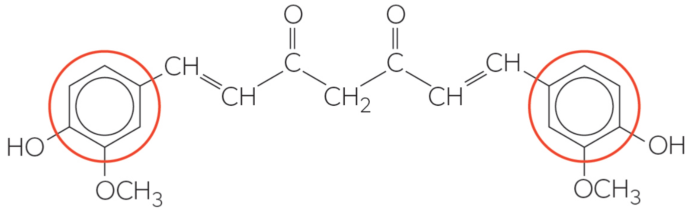
1.

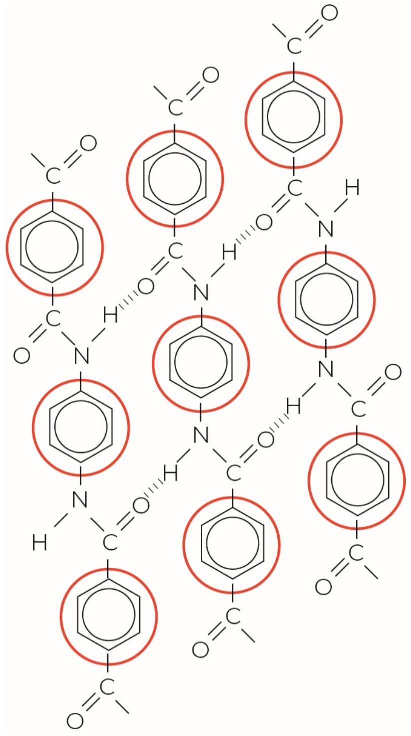


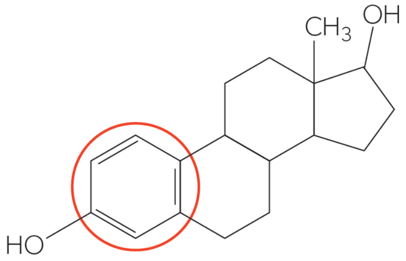
2.



3.







Practice questions (page 167)

42. Physical evidence:

- X-ray diffraction patterns show identical carbon–carbon bond lengths in benzene
- Bond enthalpy data shows that all the carbon–carbon bonds have identical strengths

Chemical evidence:

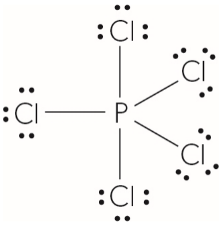
- Benzene does not easily undergo addition but undergoes electrophilic substitution instead
- The hydrogenation of benzene is less exothermic than expected
- Absence of isomers of 1,2-disubstituted benzene derivatives

43. Each carbon contributes one delocalized electron

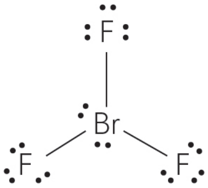
44. CCC bond angle: 120°. Molecular geometry: trigonal planar

Practice questions (page 168)

45. a.



b.



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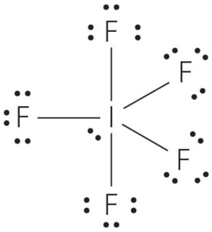

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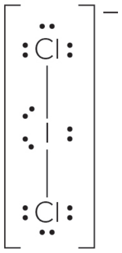
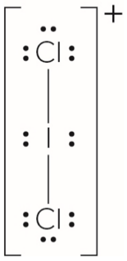
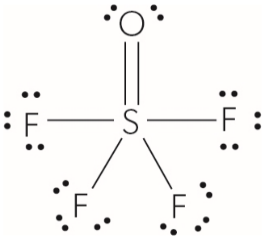


e.



f.



Practice question (page 170)

46. Equatorial domains are at 120° from each other, whereas axial domains are at 90° from other domains. Lone pairs repel more than bonding domains, so they are likely to occur in equatorial positions, where they can be further away from other domains.

Activity (page 171)

1. **a.** EDG: octahedral; MG: octahedral
   b. EDG: trigonal bipyramidal; MG: see-saw
   c. EDG: trigonal bipyramidal; MG: trigonal bipyramidal
   d. EDG: trigonal bipyramidal; MG: linear
   e. EDG: octahedral; MG: square planar
2. **a.** trigonal bipyramidal: 5 bonding domains and 0 non-bonding domains
   b. T-shaped: 3 bonding domains and 2 non-bonding domains
   c. octahedral: 6 bonding domains and 0 non-bonding domains
   d. square planar: 4 bonding domains and 2 non-bonding domains
   e. see-saw: 4 bonding domains and 1 non-bonding domain
   f. square pyramidal: 5 bonding domains and 1 non-bonding domain
   g. linear: EITHER 2 bonding domains and 3 non-bonding domains **OR** 2 bonding domains and 0 non-bonding domains
3. **a.** EDG: trigonal bipyramidal; MG: trigonal bipyramidal
   b. EDG: trigonal bipyramidal; MG: T-shaped
   c. EDG: octahedral; MG: square pyramidal
   d. EDG: trigonal pyramidal; MG: linear
   e. EDG: tetrahedral; MG: bent
   f. EDG: trigonal bipyramidal ; MG: (distorted) trigonal bipyramidal
   g. EDG: octahedral; MG: octahedral
   h. EDG: trigonal bipyramidal; MG: see-saw
   k. EDG: trigonal bipyramidal; MG: see-saw
4. **a.** 90° and 120°
   b. a little under 90° (e.g. 88°)
   c. a little under 90° (e.g. 88°)
   d. 180°
   e. a little under 109.5° (e.g. 104.5°)
   f. approximately 90° and 120° (the actual angles will be slightly different because the double S=O bond repels S–F bonds stronger than S–F bonds repel one another)
   g. 90°
   h. A little under 120° and a little under 90° and 180°
   k. A little under 120° and a little under 90° and 180°

Practice questions (pages 173–174)

47. **a.** FC(H) = 0; FC(O) = 0
    b. FC(C) = 0; FC(O) = 0
    c. FC(H) = 0; FC(O) = –1
    d. FC(N) = +1; FC(O doubly bonded) = 0; FC(O singly bonded) = –1
    e. FC(S) = 0; FC(F) = 0
    f. FC(B) = 0; FC(H) = 0

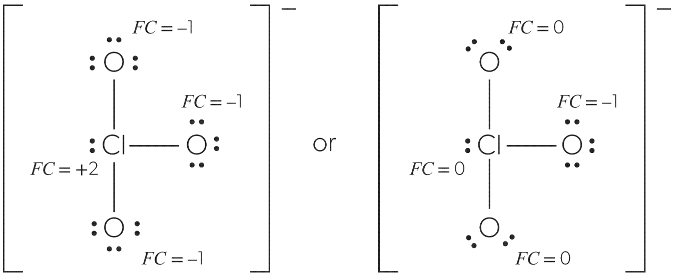
48. a. First structure

b. Third structure

49. $FC(\text{Xe}) = +3$ and $FC(\text{O}) = -1$

An alternative Lewis structure has an expanded octet on the xenon: three xenon–oxygen double bonds and a lone pair of electrons on the Xe. In that case, $FC(\text{Xe}) = 0$ and $FC(\text{O}) = 0$, which is preferable.

50.



The second (RHS) structure is preferable.

Activity (page 176)

a. 7 sigma, 1 pi

b. 6 sigma, 2 pi

c. 5 sigma, 2 pi

d. 3 sigma, 1 pi

e. 4 sigma, 0 pi

f. 3 sigma, 0 pi

g. 2 sigma, 2 pi

h. 3 sigma, 1 pi

Practice questions (page 176)

51. Sigma, σ, bonds:

- occur due to orbital overlap along the internuclear axis
- involve s–s overlap, p–p overlap, s–p overlap, etc.
- only one σ bond can be formed between any two given atoms
- involve one area of overlap per bond

Pi, π, bonds:

- occur due to orbital overlap above and below the internuclear axis
- involve p–p overlap
- one or two π bonds can be formed between any two given atoms
- involve two areas of overlap per bond
- can give rise to delocalization, if there are several alternating π bonds

Both σ and π bonds:

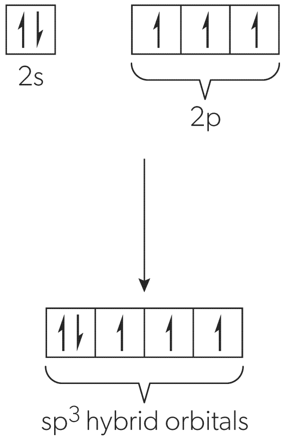
- Involve the overlap of orbitals

52. They have only one lobe each, so they cannot form two separate overlapping areas.

Practice questions (page 179)

53. See figure 88.

54.



Practice questions (page 180)

55. a. sp^3

b. sp^2

c. sp^3

d. sp

e. sp^3

f. sp^3 (if it has 4 domains: 2 bonding and 2 non-bonding) or sp^2 (if it has 3 domains: 2 bonding and 1 non-bonding)

56. Carbon in CF_4

a. number of electron domains: 4

b. electron domain geometry: tetrahedral

c. hybridization: sp^3

Carbon in HCN

a. number of electron domains: 2

b. electron domain geometry: linear

c. hybridization: sp

Nitrogen in HCN

a. number of electron domains: 2

b. electron domain geometry: linear

c. hybridization: sp

Nitrogen in N_2H_2

a. number of electron domains: 3

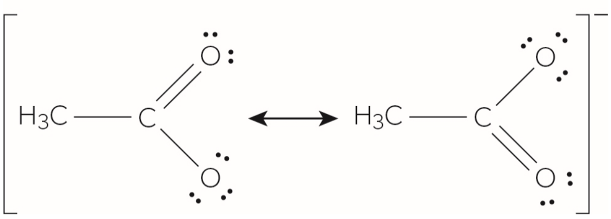
b. electron domain geometry: trigonal planar

c. hybridization: sp^2

57.

Atom	Hybridization	Number of electron domains	Electron domain geometry	Molecular geometry
A	sp^3	4	tetrahedral	tetrahedral
B	sp^2	3	trigonal planar	trigonal planar
C	sp^3	4	tetrahedral	bent

Activity (page 180)



End of topic questions (pages 181–185)

1. Answers will depend on the student, but a possible answer might be: The bonds formed between atoms that have similar electronegativity values are usually covalent, provided their mean electronegativity is high. Covalent bonding is the electrostatic attraction between a shared pair of electrons and the nuclei on either side of this shared pair. The extent to which the electron(s) in the bond are shared unequally gives rise to bond polarity.

Covalent bonds vary in their polarity. Bond polarity (or lack of thereof) contributes to the overall polarity of a molecule and impacts the substance's chemical and physical properties. For example, non-polar substances have limited aqueous solubility. Covalent compounds are typically poor electrical conductors. Their volatility and solubility vary immensely, partly due to the possibility of forming molecules.

Only covalent bonds (not ionic nor metallic) give rise to molecular substances. Many different molecules exist, each with its own properties. The behaviour of molecular substances is in great part dictated by their intermolecular forces.

2. Ionic bonds, covalent bonds and intermolecular forces are all electrostatic attractions. Ionic bonds are formed between elements that have a large difference in electronegativity. Smaller electronegativity differences give rise to covalent bonds.

Both covalent bonds and ionic bonds are chemical bonds, whereas intermolecular forces are too weak to be considered bonds as they are weak forces. Intermolecular forces are a feature of molecular substances, which are one of the two possible covalent structures.

3. Covalent structures can be network or molecular. In both cases atoms are bonded together by covalent bonds. Covalent network structures consist of continuous arrangements of atoms forming a giant lattice. They are typically insoluble in water, poor electrical conductors, and have very high melting points and boiling points. Melting and boiling require breaking many strong covalent bonds, a process which requires a large amount of thermal energy.

In contrast, molecular substances exhibit greater variation in solubility and melting point and boiling point data than covalent networks. Molecular substances are made up of discrete groups of atoms bonded together covalently. However the molecules are held close to other molecules not by bonds but by intermolecular forces. As a result, melting and boiling points of molecular substances are much lower than that of their network counterparts. This is because melting and boiling take place when these weak intermolecular forces are overcome, which does not require a large amount of thermal energy.

4. D

D is molecular because it contains bonds between nitrogen and oxygen. The electronegativity difference is low (0.4) and therefore the bonds are covalent. The other three compounds are ionic and therefore, not molecular. We know that it is a molecular compound, not a covalent network, because it does not correspond to any of the species that form giant covalent structures.

5. D

D (carbon monoxide) has the shortest carbon-to-oxygen bond because it contains a triple covalent bond. A (ethanol) and B (methoxymethane) contain single carbon–oxygen bonds, whereas C (propanone) contains a double carbon–oxygen bond.

6. C

Phosphorus trichloride has four domains: three bonding and one non-bonding. The electron domain geometry is therefore tetrahedral and the molecular geometry is trigonal pyramidal. Tetrahedral molecules have a 109.5° bond angle, but in trigonal pyramidal molecules this angle is reduced due to the repulsion exerted by the lone pair electrons on the bonding electrons.

7. A

The oxygen donates two of its lone pair electrons to form a bond with the hydrogen ion. Thus, a coordination bond is formed.

8. B

The differences in electronegativity are as follows:

C–H: $\Delta\chi = 2.6 - 2.2 = 0.4$

O–H: $\Delta\chi = 3.4 - 2.2 = 1.2$

H–Cl: $\Delta\chi = 3.2 - 2.2 = 1.0$

C–O: $\Delta\chi = 3.4 - 2.6 = 0.8$

9. C

Buckminsterfullerene (C_{60}) is the molecular substance. A, B and C have covalent network structures.

10. B

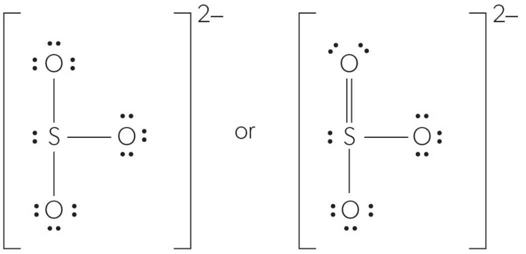
All molecules form London (dispersion) forces. In addition, CH_3F molecules are polar and therefore experience dipole–dipole forces. CH_3F does not contain a hydrogen directly bonded to a N, O or F, and therefore it does not experience hydrogen bonding between its molecules.

11. C

The important thing to realize is that there are two types of oxygen atom in the formula: the three that are single bonded to the phosphorus and one that is double bonded to the phosphorus. FC(P) = 0; FC(O doubly bonded) = 0; FC(O singly bonded) = –1

12. C

Two possible Lewis formulas of SO_3^{2-} are shown below. Both show four electron domains around the central sulfur atom, therefore the electron domain geometry is tetrahedral. (The molecular geometry is trigonal pyramidal.)



(Note: the second Lewis formula is preferred in terms of formal charge)

13. D

The bromine atom in BrF_5 has six electron domains, of which one is non-bonding. The molecular geometry is therefore square pyramidal. (The electron-domain geometry is octahedral.)

14. D

There are six electron domains around the central phosphorus atom, so the molecular geometry is octahedral.

15. D

D is octahedral, and, with all bonds being identical, the bond polarity is equal in all 6 domains. The electrons are distributed symmetrically across the molecule.

16. B

Only I and III result in sigma bonds because they involve orbitals overlapping along the internuclear axis. The orbitals in II would overlap above and below the internuclear axis, thus forming a pi bond, not a sigma bond.

17. C

The oxygen atom in ethanol has four domains, and is therefore sp^3 hybridized.

18. A

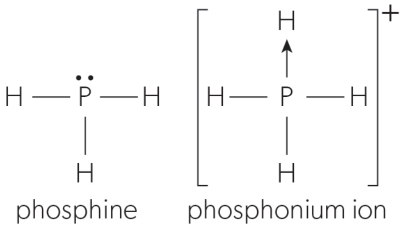
The carbon atom in HCN has two domains and is therefore sp hybridized.

19. C

The molecule contains 2 sigma bonds and 2 pi bonds. The hydrogen–carbon bond is a single bond. The carbon–nitrogen bond is a triple bond. The single bond is composed of a sigma bond. The triple bond is composed of one sigma bond and two pi bonds.

20. The electrostatic attraction between a shared pair of electrons and atomic nuclei.

21. a.



b. Phosphine: EDG: tetrahedral; MG: trigonal pyramidal

Phosphonium ion: EDG: tetrahedral; MG: tetrahedral

c. Phosphine bond angle <109.5° (the theoretical value is ~93°)

Phosphonium ion bond angle = 109.5°

d. Phosphine: polar. The electron density is unevenly distributed across the molecule, with higher electron density in the lone pair.

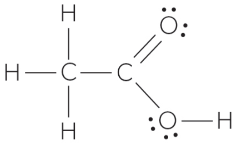
Phosphonium ion: non-polar. The electron density is evenly distributed across the molecule.

22. The three species have four electron domains. In methane all four domains are bonding domains and therefore they exert equal repulsion, making the bond angles 109.5°. In ammonia one of the domains is a lone pair of electrons. Lone pairs exert greater repulsion than bonding pairs. Therefore the bond angle decreases as the bonding domains are pushed closer together. The same applies to water, except that in water two non-bonding domains are present on the central atom. Two lone pairs exert greater repulsion than just one, pushing the bonding domains even closer together in water.

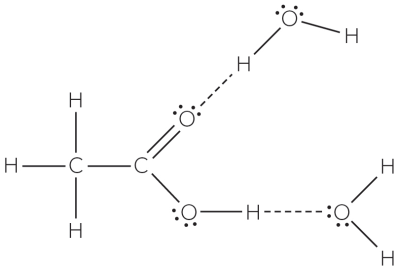
23. In graphite one electron per carbon atom is delocalized. The carbon atoms are sp^2 hybridized, and the remaining p electrons are delocalized above and below the plane of each of the sheets in graphite. Graphite is a good electrical conductor due to the presence of mobile charged particles (i.e. the delocalized electrons).

In diamond, the carbon atoms are sp^3 hybridized. There are no delocalized electrons or other mobile charged particles, so diamond is a poor electrical conductor.

24. a.



b. Various interactions are possible. Two possibilities are shown here:



The hydrogen bonds in the diagram must be clearly different from the covalent bonds.

c. i. (C × 4) + (O × 4) + (H × 8) = (12.01 × 4) + (16.00 × 4) + (1.01 × 8) = 120.12 g mol^{-1}

ii. Ethanoic acid is polar and thus associates with polar solvents through dipole–dipole interactions. If the solvent is protic, hydrogen bonds can also be formed. Being polar, ethanoic acid does not associate well with non-polar substances. In non-polar solvents, ethanoic acid molecules are better able to hydrogen bond with another ethanoic acid molecule.

d. All carboxylic acids can form hydrogen bonds with water molecules. Compared with ethanoic acid, hexanoic acid contains a longer non-polar hydrocarbon chain, which does not readily associate with polar solvents such as water. The aqueous solubility of hexanoic acid is low because the hydrocarbon chain makes up a larger proportion of the molecule than in ethanoic acid.

e. sigma: 7, pi: 1

f. C in CH_3: sp^3; C in C=O: sp^2; O in C=O: sp^2; O in O–H: sp^3

25. a. From lowest to highest boiling point:

butane, methoxyethane, propan-1-ol, ethanoic acid

Both propan-1-ol and ethanoic acid form hydrogen bonds (as well as dipole–dipole interactions and London dispersion forces). Hydrogen bonds are a strong type of IMF and therefore overcoming them requires a large amount of thermal energy, conferring a high boiling point to these substances. Of the two, ethanoic acid has the higher boiling point because each of its molecules forms a greater number of hydrogen bonds.

Methoxyethane is polar and therefore dipole–dipole interactions form between its molecules, in addition to London dispersion forces. Its boiling point is below that of propan-1-ol and ethanoic acid.

All four substances experience London dispersion forces of similar strength because their molar masses are similar. Butane does not experience IMFs of any other type and therefore has the lowest boiling point. The LDFs holding its molecules together are relatively weak, needing a relatively low amount of thermal energy to be overcome.

b. The strength of London dispersion forces increases with increasing molar mass. Therefore, when comparing the strength of different intermolecular force types, the molar masses should be more or less constant.

26. The group 17 elements form diatomic molecules, which are non-polar. The molecules in non-polar substances are held together by weak London dispersion forces. Down the group, the molar mass increases and so does the LDF strength. The boiling point increases with increasing LDF strength.

27. a. Dipole–induced dipole forces. Water is polar, while diatomic oxygen is non-polar. The presence of polar water molecules induces a dipole in the O_2 molecules.

b. Water molecules are polar and therefore form a “hydration shell” around cations and anions. The dipoles in water molecules align with the ion, stabilizing it. Non-polar solvents do not contain dipoles and therefore cannot stabilize ions in this way.

28. a. Morphine has two hydroxyl groups which can hydrogen bond with water. Therefore, morphine has higher aqueous solubility. In diamorphine, the hydroxyl groups are substituted with less polar ester groups, which do not associate with water as readily.

b. Diamorphine contains ester groups instead of morphine's hydroxyl groups and is therefore less polar. Because non-polar substances associate well with other non-polar substances, diamorphine more readily dissolves in non-polar solvents.

29. a. $$R_F = \frac{\text{distance travelled by the spot}}{\text{distance travelled by the solvent}}$$

$$= \frac{4.10\text{ cm}}{4.90\text{ cm}}$$

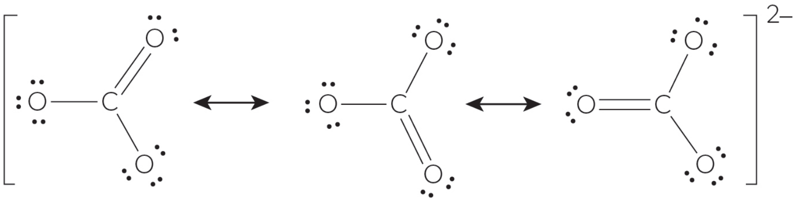
$= 0.837$ (3 sf)

b. X travelled farthest up the chromatogram, demonstrating a higher affinity for the non-polar solvent than for the polar stationary phase, Therefore, X is most likely the least polar component.

30. Conclusions may include:

- Sample S is a mixture.
- Sample S contains colourings E102 and E124 and a third which is not E110 or E122.
- Sample S does not contain colourings E110 or E122.
- The third, unidentified component is polar since it has low affinity for the solvent, which is non-polar.
- Colouring E122 is non-polar, because it has a high affinity for the solvent, which is non-polar.

31. a.



b. 4 electron pairs ÷ 3 domains = 1.33. The carbon–oxygen bond order is 1.33.

32. a. The carbon atoms in benzene are sp^2 hybridized and have trigonal planar geometry. Therefore, benzene rings are planar. Each of the six carbon atoms has an unhybridized p orbital, with lobes above and below the plane of the molecule. The electrons in these six unhybridized p orbitals become delocalized, forming a ring above and below the plane of the molecule.

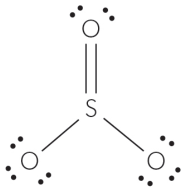
b. Possible evidence from bond length data includes:

- The carbon–carbon bond length in benzene is 140×10^{-12} m. This value lies between the carbon–carbon single bond and double bond lengths, which are 154×10^{-12} m and 134×10^{-12} m, respectively. The carbon–carbon bonds in benzene are therefore all stronger than a single bond but weaker than a double bond, suggesting bond order 1.5.
- Data show that the carbon–carbon bond lengths in benzene are all equal to each other. From this, it is inferred that the carbon–carbon bonds in benzene are all equivalent and the molecule is therefore symmetrical.

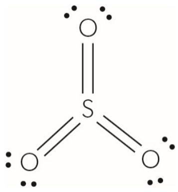
c. Possible chemical evidence includes:

- Unlike other unsaturated hydrocarbons, benzene does not readily undergo addition reactions. The delocalized electrons contribute to the stability to the molecule. Instead of addition, benzene readily undergoes electrophilic substitution reactions which allow the delocalized ring system to be present in the product.
- If the π electrons in benzene were in double bonds and not delocalized, we would expect its enthalpy change of hydrogenation to be three times that of cyclohexene. However, the hydrogenation is less exothermic than that. This is due to the stability that the delocalized electron system brings.
- 1,2-disubstituted benzene derivatives do not have isomers because in benzene the carbon–carbon bonds are all equivalent.

33. a. Lewis formula that follows the octet rule:



Lewis formula in which sulfur has an expanded octet:



The third mark is awarded for correctly drawing the non-bonding electrons on the oxygen.

b. Lewis formula that follows the octet rule:

$FC(\text{S}) = 6 - (\frac{1}{2} \times 8) = 6 - 4 = +2$

$FC(\text{O doubly bonded}) = 6 - 4 - (\frac{1}{2} \times 4) = 6 - 4 - 2 = 0$

$FC(\text{O singly bonded}) = 6 - 6 - (\frac{1}{2} \times 2) = 6 - 6 - 1 = -1$

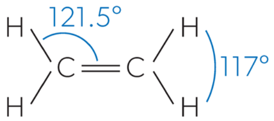
Lewis formula in which sulfur has an expanded octet:

$FC_S = 6 - (\frac{1}{2} \times 12) = 6 - 6 = 0$

$FC_O = 6 - 4 - (\frac{1}{2} \times 4) = 6 - 4 - 2 = 0$

c. The Lewis formula containing the sulfur atom with an expanded octet is preferred because the formal charges are all zero and the difference in formal charge in the molecule is also zero.

34. a.



b. trigonal planar

c. **i.** <120°, e.g. 118° (the experimental value is 117.6°)

ii. >120°, e.g. 124° (the experimental value is 121.2°)

d. sigma: 5; pi: 1

e. sp^2

f. Rotation around the carbon–carbon bond is restricted due to the presence of the pi bond above and below the plane of the molecule.

g. Ethene is a non-polar hydrocarbon. The bonds are virtually non-polar as the electronegativity difference between carbon and hydrogen is 0.4. Further, any slight bond dipoles are cancelled out due to their symmetrical arrangement in the molecule. The intermolecular force is therefore London dispersion forces.

35. Diamond: sp^3

Graphite: sp^2

Diamond does not conduct electricity because it lacks charged particles that can move across the structure: it does not contain ions or delocalized electrons. The electrons in diamond are in localized sigma bonds. Graphite does not contain ions either, but it does contain one delocalized electron per carbon atom. The delocalized electrons are p electrons, present because the carbon atoms in graphite are sp^2 hybridized.

36. **a.** Carbon: EDG: trigonal planar; MG: trigonal planar.

    Nitrogen: EDG: tetrahedral; MG: trigonal pyramidal

    b. carbon: sp^2; nitrogen: sp^3

    c. The experimental H–N–H bond angle of approximately 113° suggests a geometry somewhere between trigonal pyramidal (109.5° bond angle) and trigonal planar (120°), and a hybridization state somewhere between sp^3 and sp^2.

37. The possible configurations for a quadruple carbon–carbon bond would contain an orbital facing away from the bond.

    For example, following the pattern of bond order with regard to sigma and pi bonds, one could expect a quadruple bond to be composed of one sigma bond and three pi bonds. The formation of three pi bonds would necessitate three unhybridized p orbitals on each carbon atom to overlap on either side of the internuclear axis, but this is not possible. One of the p orbitals will have a lobe along the internuclear axis and another on the other side of the atom, facing away from the bond. In addition, there would be a very large electron density present, which would be energetically unfavourable.

Structure 2.3 – The metallic model

Skills (page 190)

Answers will vary depending on the independent and dependent variables chosen. At the very least you should identify the variables, collect data in a suitable table, plot a graph, describe and explain the trend in the graph, and any outliers (if applicable). You may also consult more than one database, attempt a semi-log plot, evaluate the data, estimate the uncertainties, etc.

The ATL task should also be introduced and planned for at the beginning of the task, to encourage the development of self-management skills.

Practice questions (page 191)

1. **a.** thermal conductivity, malleability, high melting point

   b. electrical conductivity, ductility

   c. thermal conductivity, malleability, high melting point

   d. malleability, luster, strength

   e. malleability, hardness, strength, chemical and biological resistance, etc.

   f. ductility, elasticity, sonorousness

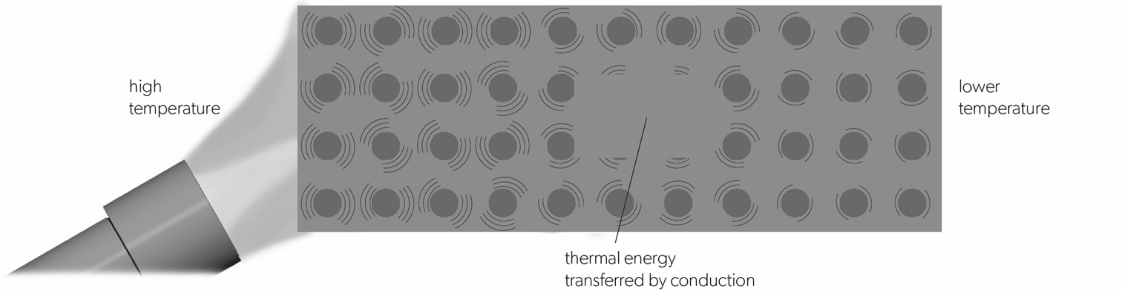
2. Diagram showing:

   - Lattice of several rows of metal cations (in a regular pattern)
   - Delocalized electrons in the spaces between the cations. See figure 3 on page 187.

3. **a.** Diagram showing that metallic bonds remain present after a row of cations is shifted due to an external force. See figure 7 on page 189.

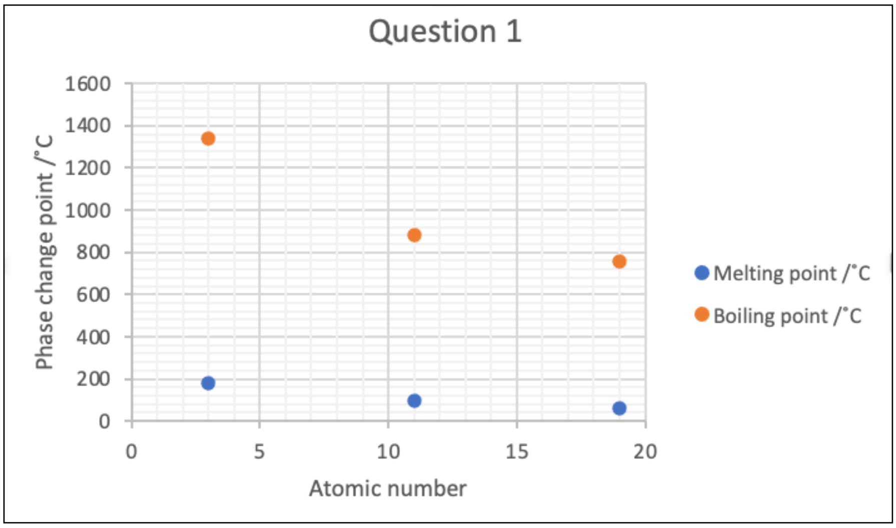
   b. Diagram showing mobility of delocalized electrons when a potential difference is applied. See figure 4 on page 188.

   c. Diagram showing electrons and ions vibrating and colliding along a piece of metal, e.g.



Skills questions (page 192)

1. For example:

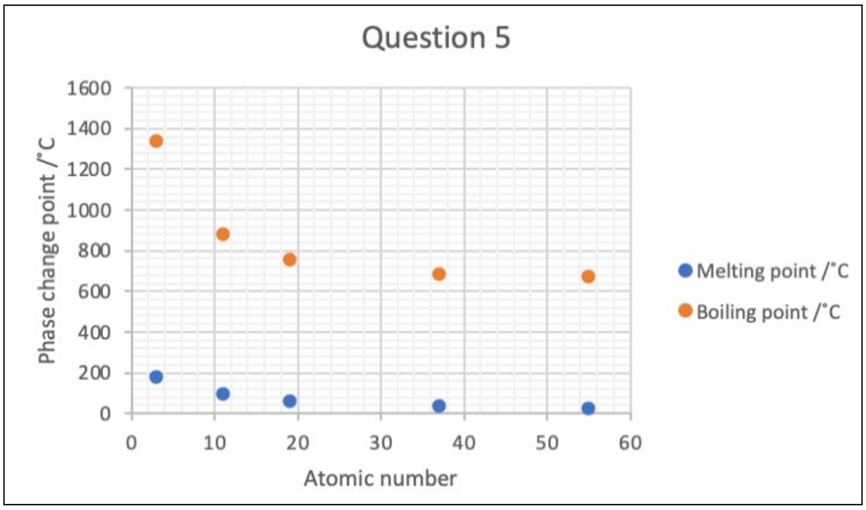


2. Answers will vary depending on the nature of the extrapolation (e.g. whether a linear relationship is assumed, whether or not the equation for the line or curve of best fit is used, etc.)

3. Answers will vary depending on the question 2 responses.

4. Answers will vary.

5. For example:



6. Answers will vary. In general, more data points lead to a better quality graph as the relationship between variables can be analysed more thoroughly.

Practice questions (page 193)

4.

- The radius of calcium ions is smaller than that of potassium ions. The electrostatic force of attraction between the positive nuclei and delocalised electrons is therefore greater in calcium.
- Calcium ions have a greater charge (+2) than potassium ions (+1). This also results in a stronger electrostatic attraction due to the greater charge density that results.
- In calcium, two electrons per atom become delocalized, whereas in potassium it is only one electron. There exists greater electron density in calcium’s “sea” of delocalized electrons, leading to a stronger metallic bond in calcium.

5. D

These metals are all in group 2. Melting point decreases down the group due to increasing ionic radius.

6. D

Electron density in the “sea” of delocalized electrons (item I), ionic radius (item II) and ionic charge (item III) all influence the electrostatic attraction between the cations and electrons, and thus the melting point.

Practice questions (page 194)

7. C

Transition elements are generally better electrical conductors than main group metals because their d electrons also become delocalized.

8. A

While transition elements do have high melting points and boiling points, this is due to the presence of delocalized d electrons, not whether they are unpaired or not.

9. Zn forms one stable ion: the Zn^{2+} ion.

Zn: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,4s^2\,3d^{10}$

Zn^{2+}: $1s^2\,2s^2\,2p^6\,3s^2\,3p^6\,3d^{10}$

Both Zn and Zn^{2+} have full d subshells. Zn is therefore not a transition element, as transition elements are defined as those that have atoms, or give rise to ions, with incomplete d sublevels.

End of topic questions (pages 195–196)

1. Answers will depend on the student, but a possible answer might be: The bonds formed between atoms that have similar electronegativity values are usually metallic, provided their mean electronegativity is low. Metallic bonding is the electrostatic attraction between delocalised electrons and the lattice of positive ions. The strength of a metallic bond depends on the strength of this electrostatic attraction. Factors affecting the strength of a metallic bond include the charge of the ion and ionic radius.

Metallic structures are typically good electrical conductors (due to the mobility of delocalised electrons within the structure), with high melting points and boiling points (due to the strong electrostatic attractions).

2. Ionic bonds, covalent bonds and metallic bonds are all electrostatic attractions. Ionic bonds are formed between elements that have a large difference in electronegativity. Smaller electronegativity differences give rise to covalent bonds or metallic bonding, depending on the mean electronegativity of the elements involved in the bond. These chemical bonds are all very strong and are all much stronger than intermolecular forces.

The difference between ionic, covalent and metallic bonding depends on the positive and negative species involved in the electrostatic attraction. Covalent bonds are the electrostatic attraction between a shared electron pair and atomic nuclei. Ionic bonds are electrostatic attractions between oppositely charged ions. Metallic bonds result from the electrostatic attraction between delocalised electrons and a lattice of positive ions.

Metallic and ionic bonding are non-directional. In covalent bonds, electrons are localised between the atoms involved in the bond.

3. The three types of bonds can lead to the formation of giant lattices. In addition, covalent structures can also be molecular. Giant lattices have high melting points and boiling points because of the strong electrostatic forces of attraction that hold their component atoms or ions together. In contrast, molecular substances have low melting points and boiling points. This is because melting and boiling happen when the molecules have enough energy to overcome the weak intermolecular forces that otherwise hold them close to other molecules.

4. A

B is incorrect because metals form cations, not anions. C is not the best answer because it is not specific: neither “delocalized” nor “cations” are mentioned. D is incorrect because it describes ionic bonding.

5. D

Metals are good electrical conductors when solid. They also have high melting points, are insoluble in water and can conduct electricity when molten.

6. a. Metallic bonding is non-directional. The delocalized electrons allow layers of cations to slide past each other, rearranging the shape of the lattice without breaking the electrostatic attraction between cations and electrons.

   Diagram – see figure 7 on page 189.

   b. i. The electrical conductivity of the metals in period 3 increases across the period from Na to Al. This is because the number of electrons contributed by each atom to the "sea" of delocalized electrons increases across the period. Greater electron density results in greater electrical conductivity. In addition, the charge density of the cations increases across the period due to decreasing ionic radius and increasing charge.

   ii. At higher temperatures cations vibrate more. Electrons collide more frequently with the vibrating cations, interrupting their path as they move through the structure. This increases the resistance to the flow of electrical current, lowering the conductivity.

   iii. Copper is a transition element. In addition to its valence electrons, it contains delocalized d electrons and therefore has greater electrical conductivity than the main group elements shown.

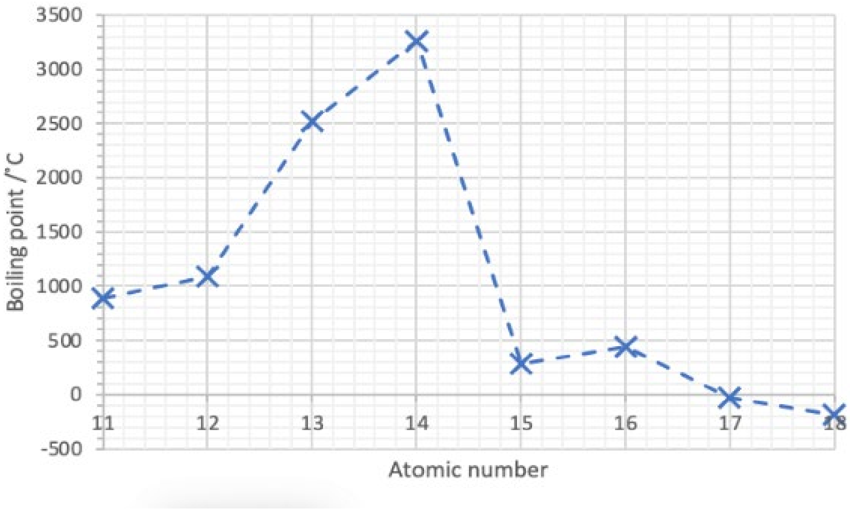
7. a. Na, Mg and Al: metallic structures

   Si: covalent network structure

   P, S and Cl: molecular covalent

   Ar: monatomic

   b.



   c. Boiling point increases from Na to Al because the strength of the metallic bond increases across the period. This is because of the greater electron density in the "sea" of electrons. In addition, the charge density of the cations also increases with increasing charge of the metallic cation and decreasing ionic radius.

   Silicon has the highest boiling point of the period 3 elements, by far. This is because overcoming the strong covalent bonds between the Si atoms in the network requires a large amount of thermal energy.

   The boiling points of non-polar molecules of P_4, S_8 and Cl_2 are low, particularly for Cl_2 as it has a relatively low molecular mass and therefore weak London dispersion forces between its molecules. S_8 has a slightly higher boiling point than P_4 due to its larger molecular mass and therefore stronger London dispersion forces.

   The boiling point of Ar is very low because it experiences only weak London dispersion forces between its atoms.

d. Na, Mg and Al are good electrical conductors because they contain delocalized electrons. Electrical conductivity increases from Na to Al because the electron density of the sea of delocalized electrons increases across the period.

Si atoms bond covalently in a covalent network structure. Si is a semiconductor. Its electrical conductivity is fairly low at room temperature, because of the lack of mobile charged particles. (As with other semiconductors, the electrical conductivity of Si increases upon heating, as electrons move to a region known as the conduction band, but this is beyond the scope of this course.)

P, S and Cl each bond covalently in molecules (most commonly P_4, S_8 and Cl_2). Molecular structures are poor electrical conductors because they do not contain mobile charged particles: their electrons are all "locked up" and localized in covalent bonds. (Note that one of the less-common allotropes of phosphorus can in fact conduct electricity.)

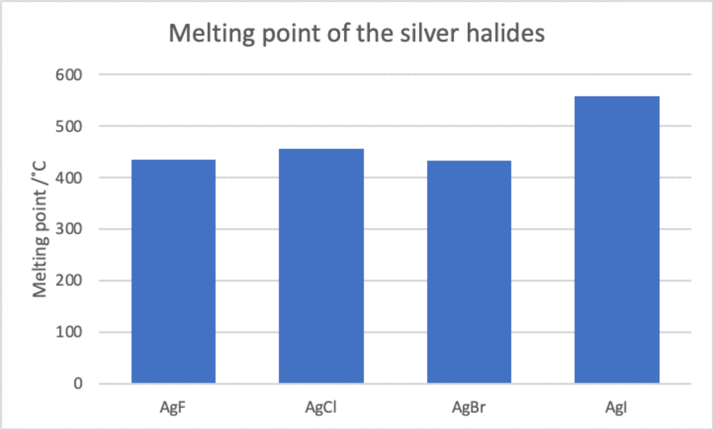
Ar is a noble gas. As such, it is found in monatomic form. It has low electrical conductivity due to the absence of mobile charged particles.

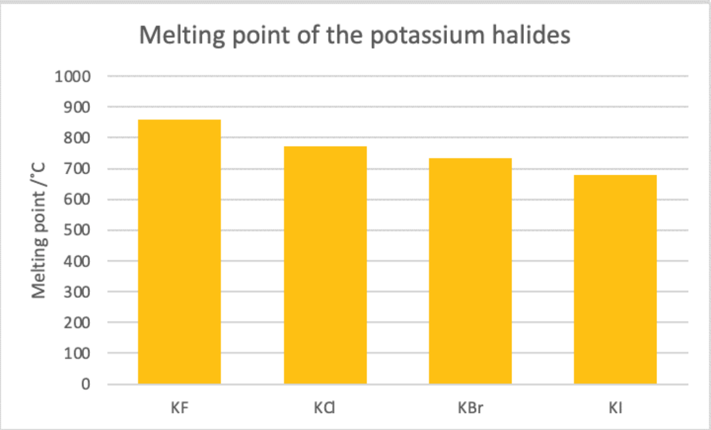
8. For a metal to change state from solid to liquid to gas, the electrostatic attractions between cations and delocalized electrons must be overcome. The stronger these forces of attraction, the greater the amount of thermal energy needed to overcome them. Therefore, melting point and boiling point serve as indicators of metallic bond strength.

Structure 2.4 – From models to materials

Data-based questions (page 200)

Answers will vary. Example bar graphs are shown here.





Answers to the second part of the question will also vary, but in general, graphs allow us to visualise a trend more easily.

Practice questions (page 203)

1. **a.** $\Delta\chi = 4.0 - 4.0 = 0$

$$\bar{\chi} = \frac{(4.0 + 4.0)}{2} = 4.0$$

The fluorine–fluorine bond is covalent.

b. $\Delta\chi = 4.0 - 0.8 = 3.2$

$$\bar{\chi} = \frac{(4.0 + 0.8)}{2} = 2.4$$

The caesium–fluorine bond is ionic.

c. $\Delta\chi = 3.4 - 1.0 = 2.4$

$$\bar{\chi} = \frac{(3.4 + 1.0)}{2} = 2.2$$

The lithium–oxygen bond is ionic.

d. $\Delta\chi = 1.9 - 1.6 = 0.3$

$$\bar{\chi} = \frac{(1.9 + 1.6)}{2} = 1.75 \approx 1.8$$

The copper–zinc bond is metallic.

e. $\Delta\chi = 3.2 - 1.6 = 1.6$

$$\bar{\chi} = \frac{(3.2 + 1.6)}{2} = 2.4$$

The aluminium–chlorine bond is found in the ionic region of the diagram, close to the polar–covalent region; this bond has a significant covalent contribution (around 50%).

f. $\Delta\chi = 3.4 - 1.9 = 1.5$

$$\bar{\chi} = \frac{(3.4 + 1.9)}{2} = 2.65 \approx 2.7$$

The silicon–oxygen bond is polar covalent. It has a significant ionic contribution.

2. **a.** The fluorine–fluorine bond is purely covalent, with zero ionic or metallic contribution.

b. The caesium–fluorine bond is highly ionic, with a very small covalent contribution and little to no metallic contribution.

c. The lithium–oxygen bond is primarily ionic, with low covalent and metallic contributions.

d. The copper–zinc bond is primarily metallic, with little to no ionic contribution. It is located close to the metallic–covalent boundary and may therefore exhibit some covalent contribution.

e. The aluminium–chlorine bond is primarily ionic, although with significant covalent contribution. Little metallic contribution.

f. The silicon–oxygen bond is primarily covalent, with some ionic contribution, making it polar covalent. Little metallic contribution.

3. The parameters only consider the electronegativity values of the two atoms involved, not the bond order or the molecular formulas. Therefore NO and NO_2 have the same relative contributions of different bonding types, according to the triangular bonding diagram.

Data-based questions (page 204)

NaCl: 5.33% error; $\Delta H_{\text{lattice}}$ = 40 kJ mol^{-1}

The experimental and theoretical values are in close agreement, suggesting a high ionic character.

AgCl: 25.1% error; $\Delta H_{\text{lattice}}$ = 184 kJ mol^{-1}

The experimental value is much higher than the theoretical value, suggesting a significant covalent contribution.

Activity (page 205)

The hydrogen–oxygen bond is covalent and highly polar; hence its ability to form intermolecular hydrogen bonds. One would expect to find it in the upper section of the covalent region in the triangular bonding diagram.

$\Delta\chi = 3.4 - 2.2 = 1.2$

$$\bar{\chi}_p = \frac{(3.4 + 2.2)}{2} = 2.8$$

The hydrogen–oxygen bond is in the polar covalent region of the diagram, as expected. Compounds containing this bond, such as water exhibit properties of polar covalent molecules. As is typical of polar molecules, water is a poor electrical conductor. It associates well with ions and is therefore a good solvent for many ionic compounds.

Practice questions (page 205)

4. **a.** The bond would be polar covalent. This means that it would have some covalent properties and also exhibit certain ionic properties. These can vary widely. For example, a compound containing this bond might have low electrical conductivity. Its boiling and melting points will likely be higher than those of covalent compounds that do not contain polar bonds. Examples of compounds roughly found in this region of the bonding triangle include SiO_2, which has a covalent network structure.

b. This substance is in the covalent region of the triangular bonding diagram, close to the boundary with the metallic region. It may well be like silicon, a semiconductive material with a mixture of metallic and covalent properties. For example, it may be lustrous and yet brittle.

c. This substance is found in the ionic region of the triangular bonding diagram. Its properties will likely be typically ionic: for example, good electrical conductivity when molten but poor when solid, crystalline and brittle.

Skills questions (page 207)

1. An indenter is pushed into the surface of sample of the metal with a known force. The dimensions of the indentation can be used to deduce the Vickers hardness value.

2. Answers will vary but they should include the four elements indicated in the question. For example: *How does the carbon content (percentage by weight) of a medium carbon steel alloy affect the alloy's hardness value, measured using the Vickers hardness test, using data reported by Calik* et al. *(2010)?*

3. The points suggest a positive relationship between carbon content and hardness. The line of best fit suggests a linear relationship.

4. Answers will vary. For example, the relationship might be non-linear and a curve of best fit is advised in that case.

5. Answers will vary.

6. Answers will vary.

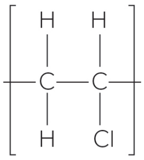
Practice questions (page 209)

5. Many alloys contain impurities whose atoms differ in size to the original metal's. The presence of impurities in alloys thus prevents the layers of atoms/ions from sliding past each other smoothly.

6. Alloys often contain component atoms that differ in size. This prevents electrons from flowing in a direct path, increasing the resistance of the material. The extent of this depends on the additional electrons provided by the impurity. If the impurity can contribute a large number of delocalized electrons, its presence may also offset some the increased resistance.

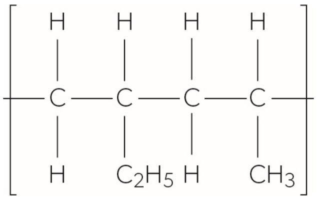
Practice questions (page 210)

7. a.

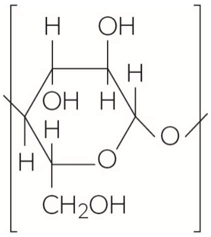


Note: in this and following answers, the square brackets around a single repeating unit are optional.

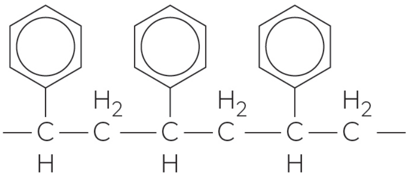
b.



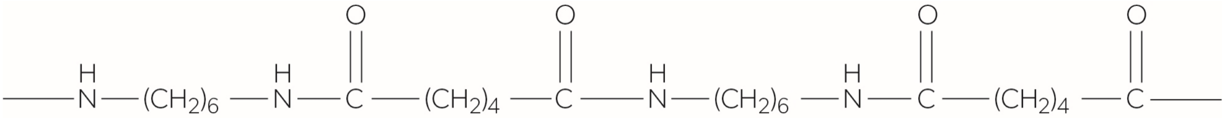
c.



8.



9.



Practice questions (page 212)

10. Polymers are long macromolecules. Their large molecular mass gives rise to strong London dispersion forces. Their length allows for a large surface through which the molecules can associate with others through intermolecular forces. The third reason their melting points are high is that separating the long, entangled molecules requires energy.

11. Polymers are covalent structures. They are electrical insulators because they lack mobile charged particles. Their electrons are localized in covalent bonds.

Activity (page 213)

Answers will vary, but students should consider whether the two polymers chosen are natural or synthetic, addition or condensation polymers, and their properties and uses. Similarities will include their long chains, covalent macromolecules, poor electrical conductivity, high specific strength, low chemical reactivity, lack of aqueous solubility, durability, etc. Differences include their uses, identity of the monomer, source (natural or synthetic), type (addition or condensation), etc.

Skills questions (page 214)

Answers will vary depending on the nature of the solutions and contact lenses used.

Practice questions (page 216)

12 Answers may include:

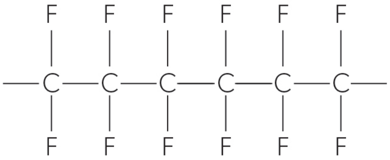
- Recycling degrades the quality of plastic materials
- Need for prior sorting and cleaning
- Energy-intensive process because the materials need to broken up and melted
- Release of atmospheric pollutants in some cases

13. Answers may include:

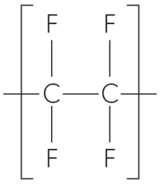
- Increased surface area to facilitate bacterial growth
- Functional groups that are attractive to bacteria
- Functional groups that are light-sensitive
- Shorter polymer chains

Practice questions (page 219)

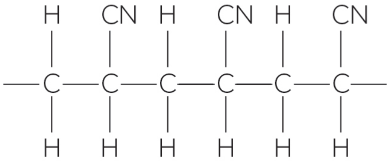
14. a. i.



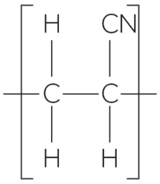
ii.



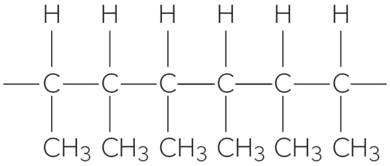
b. i.



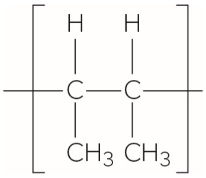
ii.



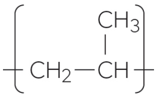
c. i.



ii.

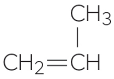


15. a. i.



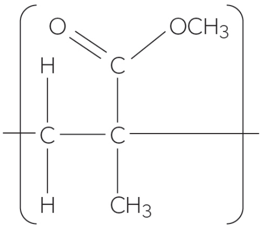
$$\left[\begin{array}{r} CH_3 \\ | \\ -CH_2-CH- \end{array}\right]$$

ii.

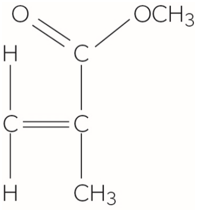


$$\begin{array}{r} CH_3 \\ | \\ CH_2{=}CH \end{array}$$

b. i.



ii.



16. Similarities:

- Both are hydrocarbons
- Both are non-polar
- Both are molecular covalent
- Both are poor electrical conductors

Differences:

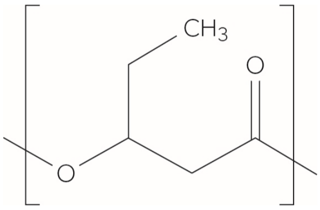
- Poly(ethene) is macromolecular and ethene is a very small molecule
- Poly(ethene) is a polymer and ethene is a monomer
- Poly(ethene) is saturated and ethene is unsaturated
- The melting point of poly(ethene) is much higher than that of ethene

Activity (page 223)

Many amide linkages are present in the image – for example, 7 of them are circled below:

Practice questions (page 223)

17.



18. Differences may include:

- Addition polymerization needs a double bond to be present. Condensation polymerization requires reactive functional groups at both ends of the monomer.
- Addition polymerization has 100% atom economy. Condensation polymerization does not.
- Condensation polymerization can be reversed through hydrolysis. Addition polymerization products are generally inert and the process is practically irreversible.
- Condensation polymerization is used to make many biomolecules while addition polymerization is used to making many synthetic polymers (plastics).

19. a. A polyester is formed, releasing H_2O

b. A polyamide is formed, releasing H_2O

c. A polyether is formed, releasing H_2O

d. A polyamide is formed, releasing HCl

End of topic questions (pages 225–227)

1. Answers will depend on the student, but a possible answer might be: The desired properties of materials depend on their intended use. The properties are, in turn, influenced by the chemical bonding and structure of the substances within. For example, electric cables need to be made from materials that are good electrical conductors as well as being able to be drawn into wires. Copper is therefore a popular choice of material for making wires. It is an excellent electrical conductor, and like all metals, it is ductile. For safety reasons, electrical wiring must be covered in an insulating material, which is why they are coated in plastic. Plastics are polymers made from covalent substances, which is why they do not conduct electricity: they do not contain charged particles capable of moving across the structure.

The properties of a substance can be enhanced by combining it with a different substance. For example, alloys. If a metal is intended for a use that requires it to be exceptionally strong, construction iron for example, it can be reinforced by alloying it with small amounts of carbon. Composites are a further example of how the properties of different substances can complement each other. Composites are heterogeneous mixtures of two phases: a reinforcing phase embedded in a matrix phase. The reinforcing phase provides structure, while the matrix phase bulks it up. As with all mixtures, the two components retain their original properties and, because of this, composites have amalgamated properties that are uniquely useful. Examples of common composites are given in the table below.

Composite	Reinforcing phase	Matrix phase	Examples of uses
concrete	aggregate	cement	construction
reinforced concrete	rebar (steel bars)	concrete	construction
fibreglass	glass fibre	plastic	boats, bathtubs, wind turbines, storage tanks, car parts
carbon fibre	carbon fibre	polymer	aircraft, sports equipment, prosthetic limbs
carbon-carbon	carbon fibre	carbon	high temperature applications in aerospace engineering, car parts
straw and clay	straw	clay	construction
wood	cellulose	lignin	construction, wooden artefacts

2. C

The mean electronegativity is 2.4 and the electronegativity difference is 1.2. When calculating these electronegativity parameters the number of each type of atom in the formula does not need to be taken into account.

3. C

ZnS is located near the boundary between ionic and covalent regions in the bonding diagram.

	Electronegativity difference	Mean electronegativity	Location in bonding diagram
A	0.1	0.85 ≈ 0.9	Metallic
B	3.2	2.4	Ionic
C	1.0	2.1	Ionic/covalent boundary
D	0.0	3.0	Covalent

4. B

B is a mixture of two metals, i.e. an alloy. A describes a solution. C is a composite. D, magnesium and chlorine would react to form magnesium chloride, which is a compound, not a mixture.

5. A

A is an alkane and is therefore saturated. B, C and D are all able to form addition polymers because they are unsaturated. Drawing Lewis formulas will also reveal the carbon–carbon double bond in these compounds.

6. D

All the pairs of molecules contain two reactive groups in each molecule that can react through condensation. The molecules in D would form an ester, but not a polymer because they each only have one reactive functional group.

7. Electronegativity values: Sn = 2.0, Pb = 1.8, Cl = 3.2

$$\overline{\chi}(SnCl_2) = \frac{(3.2 + 2.0)}{2} = 2.6$$

$$\overline{\chi}(PbCl_2) = \frac{(3.2 + 1.8)}{2} = 2.5$$

$\Delta\chi(SnCl_2) = 3.2 - 2.0 = 1.2$

$\Delta\chi(PbCl_2) = 3.2 - 1.8 = 1.4$

$SnCl_2$: Plotting (x, y) coordinates of (2.6, 1.2) on the bonding triangle diagram classifies $SnCl_2$ as mostly covalent (60–70% covalent character).

$PbCl_2$: Plotting (x, y) coordinates of (2.5, 1.4) on the bonding triangle diagram classifies $PbCl_2$ as more ionic (approximately 60% ionic, 40% covalent).

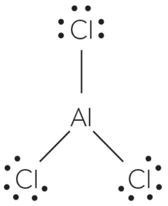
Because $SnCl_2$ shows more covalent property than does $PbCl_2$, it is more likely to exist as discrete molecules when vaporized. It will also have a lower melting point, as covalent substances have lower melting and boiling points than do ionic substances.

8. **a.** **i.** $\Delta\chi = 3.2 - 1.6 = 1.6$

$$\overline{\chi} = \frac{(3.2 + 1.6)}{2} = 2.4$$

ii. The aluminium–chlorine bond is found right around the covalent/ionic boundary in the bonding diagram. Aluminium chloride therefore has properties of both covalent and ionic substances.

b.



c. Molecular geometry is trigonal planar; bond angle is 120°

d. **i.** A covalent bond where both the electrons come from the same atom.

ii. FC(Al) = 3 – (0 + 5) = –1

FC(Cl top and bottom) = 7 – (6 + 1) = 0

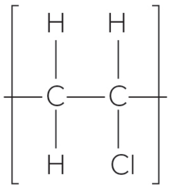
FC(Cl middle) = 7 – (4 + 2) = +1

9. The differently sized atoms/ions in alloys disrupt the regular lattice structure, interfering with the smooth movement of layers of atoms/ions over each other. See figure 8 on page 205.

10. a. A carbon–carbon double bond / it is an alkene

b. addition polymerization

c.



d. molar mass of monomer = (12.01 × 2) + (3 × 1.01) + 35.45 = 62.5 g mol^{-1}

$$\frac{69000}{62.5} = 1104 \approx 1100 \text{ (2 sf)}$$

e. Both are polar and capable of forming dipole–dipole interactions and London dispersion forces. The London dispersion forces holding chains of poly(chloroethene) together are greater due to its much larger molar mass. In addition, the long chains in poly(chloroethene) tangle up, requiring energy to separate.

11. a. PLA contains ester functional groups that can bacterial enzymes can bind to, allowing the plastic to biodegrade.

b. Possible advantages include:

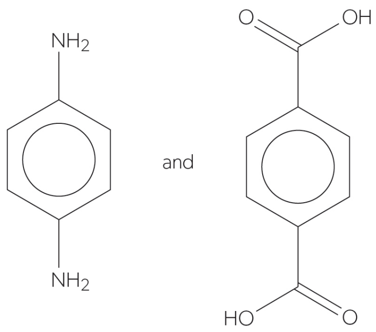
- Renewable feedstocks do not rely on decreasing amounts of fossil fuels
- Carbon-neutral
- Reduction in the volume of landfill waste
- Innocuous degradation products

Possible disadvantages include:

- Fertilizer runoff causes eutrophication
- Using land to grow feedstocks for bioplastics instead of food
- Bioplastics do not generally biodegrade under home-composting conditions; industrial composting conditions are required
- Objects made of bioplastic can biodegrade before their end of use
- Biodegradation of bioplastics in landfills can produce methane, which is a greenhouse gas

12. a. –CONH– group circled

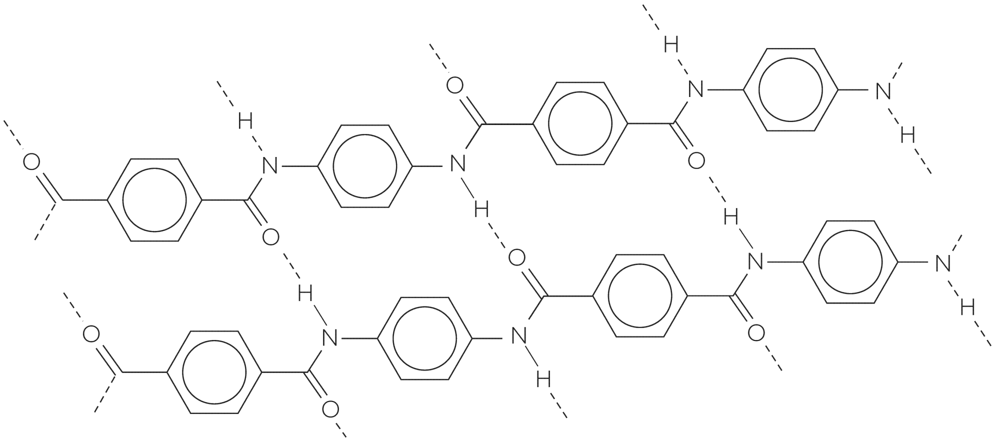
b.



c. H_2O; condensation polymerization

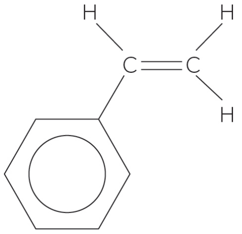
d. The hydrogen atom directly bonded to an electronegative atom (nitrogen); and the lone pairs of electrons on the oxygen.

e. A diagram showing at least two hydrogen bonds, preferably between the oxygen on one chain and the amine hydrogen on the other:

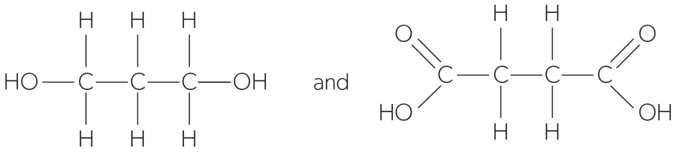


13. a. polymer A is addition; polymer B is condensation.

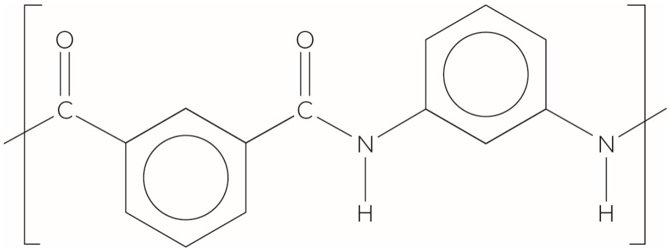
b. polymer A monomer:



polymer B monomers:



14. a.



b. HCl