Achievement Standard 91164

Demonstrate understanding of bonding, structure, properties and energy changes

CHEMISTRY

2.4

Externally assessed 5 credits

Chemical bonding and properties of molecules

This chapter assumes knowledge of atomic structure.

Chemical bonding properties

Most elements do not consist of isolated atoms but are made up of atoms held together by **chemical bonds**. A bond is an **electrostatic force of attraction** between positively and negatively charged species. There are three general ways that atoms are bonded – ionic bonding, covalent bonding and metallic bonding.

Ionic bonding

When electrons are *transferred* between atoms, electrically charged particles called **ions** are formed. The force of attraction between positive ions (cations) and negative ions (anions) is called an **ionic bond**. Monatomic cations (having only one atom) form from metal atoms, and monatomic anions are formed from non-metal atoms.

Covalent bonding

When electrons are *shared* between two atoms, a **covalent bond** is formed. The bonding electrons 'belong' to both atoms at the same time. Covalent bonds form between the atoms of non-metals.

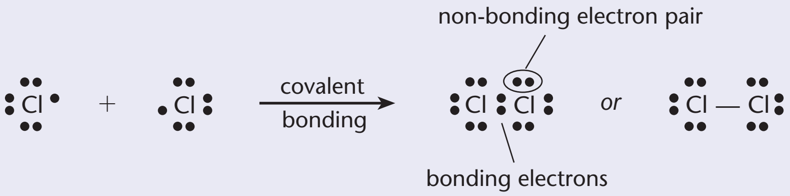
Metallic bonding

In metals, each atom is bonded to all neighbouring atoms by the electrostatic attraction between the loosely-held valence electrons and the positive metal ions (cations) that form when the valence electrons move away from the metal atom.

Lewis structures and shapes of molecules

Lewis structures (also called **electron dot diagrams**) can be used to represent the arrangement of electrons in species that contain covalent bonds.

A Lewis structure only shows the valence electrons, as these are the only ones involved in bonding. A Lewis structure must account for all the valence electrons of all the atoms in the molecule or ion.



The **shared** electrons are called **bonding** electrons. The electrons not involved in bonding are called **non-bonding** electrons or **lone pairs** of electrons. Shared electrons are usually represented by a line; e.g. chlorine, Cl_2, is Cl–Cl

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Conduction of electricity

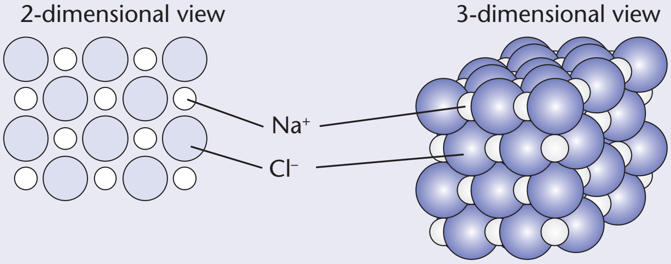
Covalent network solids usually do not consist of any free-moving charged particles so do not conduct electricity. In graphite, however, three valence electrons are involved in covalent bonding – the fourth valence electron is delocalised and free to move between the layers, so graphite *will* conduct electricity.

Solubility

Covalent network solids are insoluble in polar and non-polar solvents.

Ionic substances

Ionic substances (e.g. KCl, NaCl, MgO, MgCl$_2$) are made up of positive and negative ions in fixed positions in a three-dimensional arrangement with a regular repeating pattern known as a **lattice**. The electrostatic attraction between the oppositely charged ions (ionic bond) is strong.



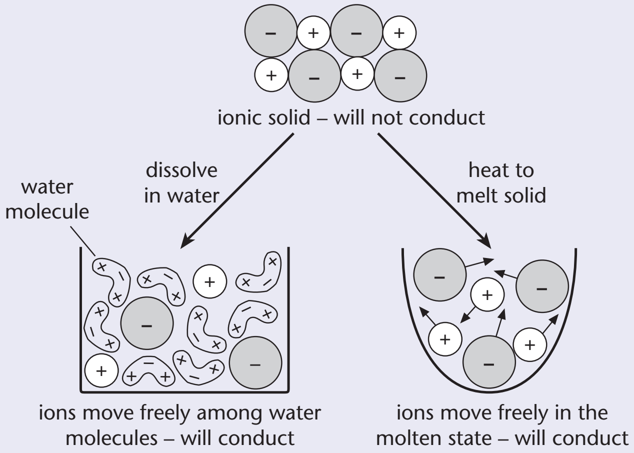
Properties of ionic substances

Melting and boiling points

The melting and boiling points of ionic substances are high because ionic bonds are strong, i.e. the positive and negative ions strongly attract one another. Considerable energy is required to separate the ions to melt or vaporise an ionic substance.

Conduction of electricity

Ions are charged particles. In the solid state, ions are held in fixed positions in the lattice and are unable to move around – the solid will *not* conduct electricity. When an ionic solid is melted or dissolved, the ions separate and are free to move around and *will* conduct electricity.



Crystal hardness

The strong electrostatic forces between the ions in the lattice produce a hard structure. Ionic solids are brittle because when pressure is applied the layers can change position so that like charges line up. These will repel each other and the solid will shatter at this point.

Solubility in polar solvents

Ionic solids dissolve in polar solvents.

Example

Sodium chloride dissolves in water

Ions are separated from the lattice when water molecules surround the ions with their negative ends pointing to Na$^+$ ions and their positive ends pointing to Cl$^-$ ions. The attraction of the water molecules for the ions is stronger than the attraction between the Na$^+$ and Cl$^-$ ions, so the salt dissolves.

Solubility in non-polar solvents

Ionic solids do *not* dissolve in non-polar solvents.

Example

Sodium chloride does not dissolve in cyclohexane

Cyclohexane molecules are non-polar, so are not strongly attracted to the Na$^+$ or Cl$^-$ ions; so, the salt will not dissolve – i.e. the Na$^+$Cl$^-$ ionic bond dominates.

Achievement Standard 91165

Demonstrate understanding of the properties of selected organic compounds

CHEMISTRY

2.5

Externally assessed 4 credits

Characteristics and physical properties of hydrocarbons

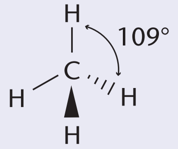
Hydrocarbons are organic compounds that contain only carbon and hydrogen. Alkanes, alkenes and alkynes are all hydrocarbons. They each form a **homologous** series (a 'family' where members of the series are represented by the same general formula, have the same functional group, have similar methods of preparation, and show similar chemical properties).

	Alkanes	Alkenes	Alkynes
General formula	C_nH_{2n+2}	C_nH_{2n}	C_nH_{2n-2}
Saturated?	yes	no	no
Functional group	–	double bond	triple bond

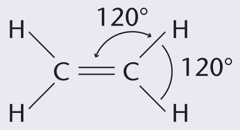
Saturated hydrocarbons contain single bonds. **Unsaturated** hydrocarbons contain a double or triple bond. A **functional group** is an atom or group of atoms within the organic molecule that determines the characteristic chemical properties of the particular family of organic compounds.

Examples of hydrocarbons

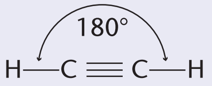
Methane is the simplest alkane and has the molecular formula CH_4 A methane molecule has a tetrahedral shape.



Ethene is the simplest alkene and has the molecular formula C_2H_4 An ethene molecule has two carbon atoms joined by a double bond. The four atoms around this double bond lie in a plane, with angles of 120° between the bonds.



Ethyne is the simplest alkyne and has the molecular formula C_2H_2 The ethyne molecule has two carbon atoms joined by a triple bond. It has a linear shape.



Physical properties of hydrocarbons

Physical property	Alkanes, alkenes and alkynes
Colour	colourless
Polarity	non-polar
Solubility in water	insoluble (ethyne is very slightly soluble)
Electrical conductivity	do not conduct
Conductivity of heat	poor
Density	less dense than water
Melting points and boiling points	• low due to weak forces of attraction between the molecules • gradually increase as molar mass increases

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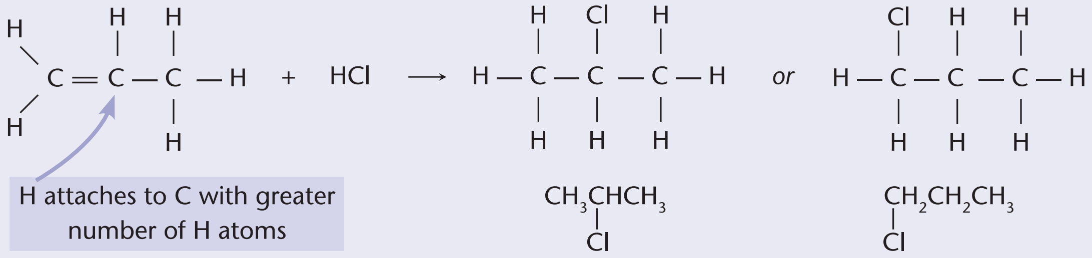
Addition of hydrogen chloride to unsymmetrical alkenes

If an alkene is **unsymmetrical**, there are two possible products from addition of hydrogen chloride. The major product has the hydrogen atom attached to that carbon atom which originally had the greater number of hydrogen atoms attached to it. ('The rich get richer.')



H attaches to C with greater number of H atoms

propene

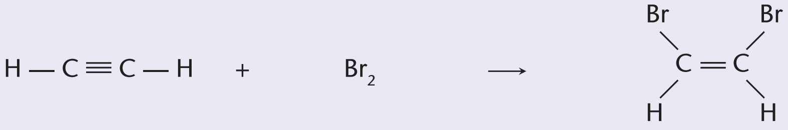
2-chloropropane
major product

1-chloropropane
minor product

The same rule applies for the addition of hydrogen bromide and water to unsymmetrical alkenes.

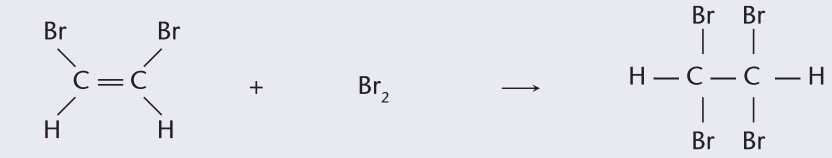
Addition reactions of alkynes

Reactions are similar to alkenes, but further addition occurs until the molecule is fully saturated.



ethyne

1,2-dibromoethene



1,2-dibromoethene

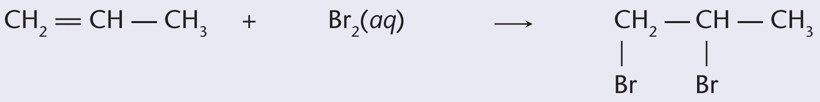
1,1,2,2-tetrabromoethane

Reactions to identify alkenes

Identification with bromine water

Alkenes

Bromine liquid is a reddish colour but, when added to water, forms orange/brown bromine water, $Br_2(aq)$. When bromine water is added to an alkene (e.g. cyclohexene – liquid at room temperature), it decolourises rapidly. Alkenes which are gases at room temperature (e.g. propene) can be bubbled through bromine water.



$$CH_2{=}CH{-}CH_3 + Br_2(aq) \longrightarrow \underset{Br}{CH_2}{-}\underset{Br}{CH}{-}CH_3$$

propene

bromine water
(orange)

1,2-dibromopropane
(colourless)

Alkanes

If an alkane, such as propane, is bubbled through bromine water, the bromine water would remain orange.

However, if placed in sunlight, the bromine water would decolourise slowly with an alkane – sunlight catalyses the reaction between alkanes and bromine.

$$CH_3{-}CH_2{-}CH_3 + Br_2(aq) \xrightarrow{UV} Br{-}CH_2{-}CH_2{-}CH_3 + HBr$$

propane | bromine water | 1-bromopropane | hydrogen bromide

Achievement Standard 91166

Demonstrate understanding of chemical reactivity

CHEMISTRY

2.6

Externally assessed 4 credits

Rates of reactions

Rate of reaction is 'the speed at which a reaction occurs'.

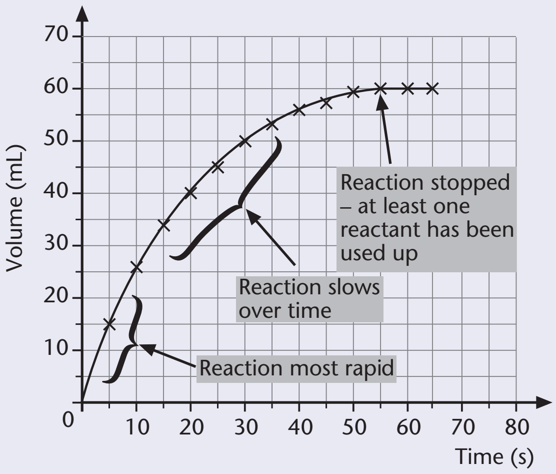
Some reactions occur quickly (e.g. adding zinc to hydrochloric acid), others occur slowly (e.g. iron rusting in a car).

Measuring rates of reactions

The rate at which a reaction occurs is measured by how quickly reactants are consumed or how quickly products form (e.g. a gas product can be measured).

The graph shows how the volume of hydrogen gas given off in the reaction between zinc and hydrochloric acid changes over time.

Volume of hydrogen gas formed against time



Collision theory

For a reaction to occur, particles must collide with sufficient energy for bonds to break and new bonds to form to make new substances.

Factors that control rates of reactions are:

- the frequency of collisions – the more collisions in a given time, the faster the reaction
- the effectiveness of the collisions – this is in terms of *orientation* (reactant particles must be positioned so that bonds break and bonds form) and *energy* (particles must collide with sufficient energy to break existing bonds and make new bonds).

Particles must collide with enough **kinetic** ('moving') energy. Total kinetic energy must be greater than the **activation energy**. Activation energy, E_A, is the minimum amount of energy required for a reaction to proceed between colliding particles.

Factors affecting rates of reactions

The rate of a reaction depends on the concentration, surface area, temperature of reactants, and the presence of catalysts.

Concentration

Concentration affects reactions in which at least one of the reactants is in solution or the reactants are gases.

With increased concentration, particles are closer together so there is an increased chance of colliding. As a result, the *frequency of collisions increases* and hence the rate of reaction increases.

The rate of a reaction will decrease as the reaction continues because the concentration of the reactants is decreasing.

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Questions: Rates of reactions

Question One

Year 2018
Ans. p. 116

a. In the iodine clock reaction, a solution of hydrogen peroxide is mixed with a solution containing potassium iodide, starch, and sodium thiosulfate.

After some time, the colourless mixture suddenly turns dark blue.

The table shows the time taken for the reaction performed at different temperatures. The concentration of all reactants was kept constant.

Temperature / °C	Time for dark blue colour to appear / s
20	15
30	9
40	4

Explain the effect of changing the temperature on the rate of reaction.

Refer to collision theory and activation energy in your answer.

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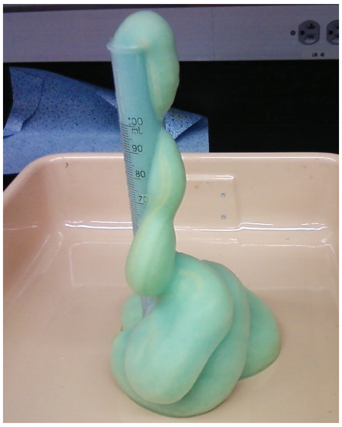
Year 2015
Ans. p. 117

Question Four



The 'elephant toothpaste' demonstration shows the decomposition of hydrogen peroxide, H_2O_2, into water and oxygen gas.

$$2H_2O_2(aq) \longrightarrow 2H_2O(\ell) + O_2(g)$$

This reaction can be observed by adding detergent to the hydrogen peroxide solution. As oxygen gas is produced, the detergent foams up, as seen in the photograph on the right. The time taken for the foam to reach the top of the measuring cylinder can be used to measure the rate of the reaction.

Three experiments were carried out to investigate factors that change the rate of the reaction.

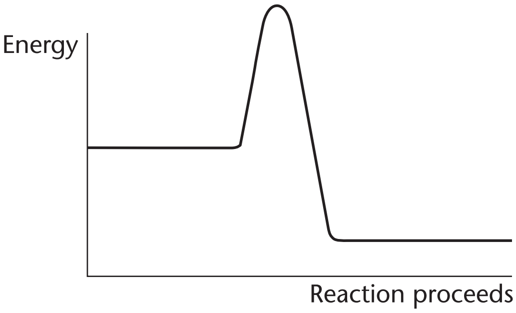
Experiment	Concentration of H_2O_2	Temperature °C	Presence of small amount of MnO_2
1	20%	20	yes
2	20%	30	yes
3	30%	20	yes

a. The decomposition reaction of hydrogen peroxide, H_2O_2, is very slow. By adding a small amount of powdered manganese dioxide, MnO_2, the rate of the reaction can be increased.

i. Explain why only a small amount of manganese dioxide is needed to increase the rate of the reaction.

ii. The diagram shows the energy diagram for the decomposition reaction without manganese dioxide.

Label this diagram and use it to help you explain how the addition of manganese dioxide speeds up the rate of the reaction.



Pressure

Pressure can be altered by changing the amount of gas in a closed system or by changing the volume of a closed system.

An increase in pressure (by adding more gas or decreasing the volume of the closed system) will be counteracted by decreasing the number of gas particles in the equilibrium mixture. Hence, the equilibrium shifts in the direction of the least number of moles of gas.

Example

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

Decreasing volume will increase pressure. The equilibrium will move to reduce the number of gas particles so will move to the right, i.e. in the forward direction (since in the reactants there is a total of 4 moles of gas molecules but there are only 2 moles in the products).

A pressure change will only have an effect if gases are present.

Addition of unreactive gases has no effect on the position of equilibrium.

Temperature

Increasing the temperature will favour the endothermic reaction. This counteracts the effect of the increase in temperature by energy being absorbed/used.

Decreasing the temperature will favour the exothermic reaction. This counteracts the effect of the decrease in temperature by energy being released.

Example

$$\underset{\text{colourless gas}}{2N_2O_5(g)} \rightleftharpoons \underset{\text{brown gas}}{4NO_2(g)} + \underset{\text{colourless gas}}{O_2(g)} + \text{energy}$$

This reaction is exothermic (ΔH is negative).

If the mixture of gases is heated, the endothermic reaction (i.e. the reverse reaction) is favoured – this will cause the amount of brown NO_2 gas to decrease. You would observe the brown colour of the reaction mixture becoming lighter. The reaction mixture will not turn colourless because, at equilibrium, all products and reactants are present, i.e. there will still be some NO_2 gas present to colour the reaction mixture brown.

In equilibrium reactions, the $\Delta_r H$ sign refers to the forward reaction. When $\Delta_r H$ is positive, the (forward) reaction is endothermic. When $\Delta_r H$ is negative, the (forward) reaction is exothermic.

Summary of responses of equilibrium system to changes

Change in conditions		Direction of change in equilibrium position
Concentration	Increase products	In the reverse direction
	Decrease products	In the forward direction
	Increase reactants	In the forward direction
	Decrease reactants	In the reverse direction
Pressure	Increase	In the direction with the least number of moles of gas
	Decrease	In the direction with the greater number of moles of gas
Temperature	Increase	In the direction of the endothermic reaction
	Decrease	In the direction of the exothermic reaction
Catalyst added		No change in equilibrium position or in K_c; equilibrium is reached more quickly (i.e. reaction rate changes).

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Questions: Equilibrium

Year 2018
Ans. p. 118

Question One

The Contact Process is used industrially in the manufacture of sulfuric acid. One step in this process is the oxidation of sulfur dioxide, $SO_2(g)$, to sulfur trioxide, $SO_3(g)$.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

a. Write the equilibrium constant expression for this reaction.

$K_c =$

b. i. Calculate the equilibrium constant (K_c) for this reaction at 600 °C using the following concentrations:

$[SO_2] = 0.100 \text{ mol L}^{-1}$

$[O_2] = 0.200 \text{ mol L}^{-1}$

$[SO_3] = 0.0930 \text{ mol L}^{-1}$

ii. Explain what the size of the K_c value indicates about the extent of the reaction at equilibrium.

c. Explain, using equilibrium principles, why it is important for an industrial plant to continue to remove the sulfur trioxide gas, $SO_3(g)$, as it is produced.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

Questions: Acids and bases

Question One

Year 2018
Ans. p. 119

a. The pH of a solution is 10.8.

i. Calculate the hydronium ion concentration, $[H_3O^+]$, and the hydroxide ion concentration, $[OH^-]$, in the solution.

$[H_3O^+]$ =

$[OH^-]$ =

ii. A sodium hydroxide solution, NaOH(*aq*), has a concentration of 0.0125 mol L^{-1}.

Calculate the pH of the sodium hydroxide solution.

iii. When the salt sodium ethanoate, CH_3COONa(*s*), is dissolved in water it separates into ions.

Use TWO relevant equations to explain whether **the solution** is acidic or basic.

b. The hydrogensulfate ion, HSO_4^-, is an amphiprotic species because it can both accept or donate a proton, thus acting as an acid or base.

Complete the equations for the reactions of the hydrogensulfate ion, HSO_4^-, with water in the box below.

HSO_4^- acting as	**Equation**
an acid	$HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons$
a base	$HSO_4^-(aq) + H_2O(\ell) \rightleftharpoons$

Answers and explanations

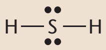
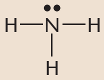
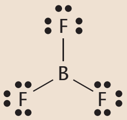
Achievement Standard 91164 (Chemistry 2.4): Demonstrate understanding of bonding, structure, properties and energy changes

2.4 Chemical bonding and molecular properties

Question One

p. 5

a.

Molecule	H_2S	NH_3	BF_3
Lewis structure			
Name of shape	Bent or V-shaped	Trigonal pyramid	Trigonal planar
Approximate bond angle around the central atom	109.5°	109.5°	120°

b. NH_3 has 4 electron clouds/regions of electron density around the central atom. To minimise repulsion, these take a **tetrahedral** shape with the angle between the electron clouds of around 109.5°. There are only 3 bonds, so the shape observed will be a **trigonal pyramid** with an angle of approximately 109.5°.

BF_3 has 3 electron clouds/regions of electrons around the central B atom. To minimise repulsion, these will take a **trigonal planar** arrangement, with the angle between the electron clouds being 120°. Since all the electron clouds are bonds, the shape of the molecules is trigonal planar.

Although both molecules have 3 bonds to the central atom, the ammonia molecule has 4 regions of electrons and the boron trifluoride has only 3 regions of electrons so the electron geometry around the central atom is different.

c. In CO_2 molecules, the bonds are polar since the electronegativity of the bonded atoms, C and O, are different. Since the shape of the molecule is linear, the polarity of the bonds is cancelled (as they are equal and opposite).

In HCN, the polarities of the C–H and the C–N bonds are different – the electron density in the C–N bond is more strongly attracted to the more electronegative N atom. This causes a charge separation across the molecule and so the molecule will be polar.

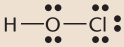
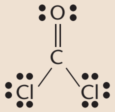
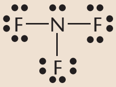
Or: The polarity of the C–N bond is greater than that of the C–H bond as the N atom is more electronegative than the H atom. Hence, the polarity of the bonds do not cancel, causing a dipole and hence a polar molecule.

(**A** – some of correct Lewis structures, correct shapes, recognises shape is related to regions of electrons on the central atom, describes polarity in relation to electronegativity; **M** – links shape/angle to number of electron clouds or links bond polarity to electronegativity and polarity of molecules; **E** – full explanations for shape and polarity)

Question Two

p. 7

a. i.

Molecule	HOCl	$COCl_2$	NF_3
Lewis structure			
Name of shape	Bent or V-shaped	Trigonal planar	Trigonal pyramid
Approximate bond angle around the central atom	109.5°	120°	109.5°

ii. Shape and bond angle are determined by the number of electron clouds / regions of electron density on the central atom. These are arranged as far apart as possible to minimise repulsion between the electron clouds. HOCl – central atom, O, has 4 electron clouds / regions of electron density around the central O atom. These take a *tetrahedral* shape with the angle between the electron clouds of around 109.5°. There are only 2 bonds, so the shape observed will be bent or V-shaped with an angle of approximately 109.5°.

$COCl_2$ has 3 electron clouds / regions of electrons on the central C atom. These will take a trigonal planar arrangement with the angle between the electron clouds being 120°. Since all the electron regions on the central atom are bonds, the shape of the molecule is also trigonal planar.

b. i. Dichloromethane – **polar**.

Tetrachloromethane – **non-polar**.

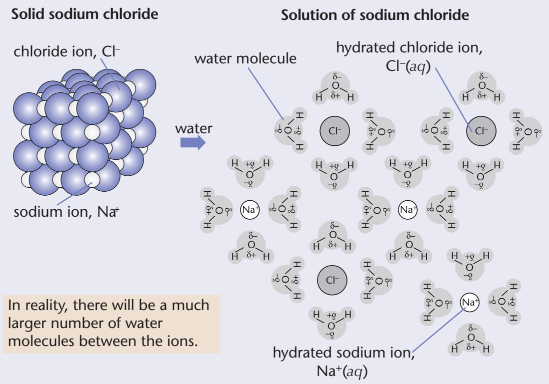
ii. CH_2Cl_2 has two different bonds – i.e. C–Cl and C–H. Each bond will be polar, since there are different atoms present; however, the C–Cl bond will be more polar since the Cl atom has greater electronegativity than the H atom. This means that the Cl atom will attract the electrons in the C–Cl bond to a greater extent than the H atom in the C–H bond will. Hence the charge separation / bond dipole in the C–Cl bond is greater than for the C–H bond. The bonds are evenly distributed around the C atom, but the different polarities mean that the electron density is not evenly spread around the C atom, the bond polarities / dipoles do not cancel and the molecule is *polar*.

In CCl_4, all the bonds are polar since the electronegativity of the C and Cl atoms is different. Since all the bonds are identical and the shape is tetrahedral, the polar bonds are evenly spread and the dipoles / bond polarities will cancel. Hence the molecule is *non-polar*.

(**A** – some of: correct Lewis structures, correct shapes, recognises shape is related to regions of electrons on the central atom, polarity of bonds or molecules; **M** – links shape/angle to number of electron clouds or links bond polarity to electronegativity or polarity of molecules to their shape; **E** – full explanations for shape and polarity)

d. **Sodium chloride dissolving in water**

$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$



Polar water molecules separate from one another as they are attracted to the charged ions in the sodium chloride lattice. The oxygen atoms are slightly negatively charged (δ–), so are attracted to the positive Na^+ ions. When a number of water molecules surround a Na^+ ion, it is able to break away from the attractive forces in the lattice and move through the water. In a similar way the hydrogen atoms in the water molecules are slightly positively charged (δ+) so are attracted to the negatively charged chloride ions, Cl^-. As the ions are attracted to the water molecules, the ionic lattice breaks up.

(**A** – *some of* identifies type of solid, particles and forces between particles, reasons for ductility, solubility, conductivity; **M** – *three or four of* links structure and bonding to properties of graphene, magnesium, iodine, explains attractions between water molecules and ions; **E** – *three or four of* explains all properties in relation to structure and bonding for graphene, magnesium, iodine, and explains solubility of NaCl with a diagram)

2.4 Energy changes

Question One

p. 28

a. **endothermic**

Reason: $\Delta H > 0$ (positive).

b. i. **exothermic**

Reason: Energy is released in the reaction.

ii. Process is endothermic – energy is needed to separate the particles in the liquid to change to the gaseous state where the particles are far apart.

c. i. $H_{\text{products}} < H_{\text{reactants}}$

Hence $\Delta H = H_{\text{products}} - H_{\text{reactants}} < 0$

Energy is lost to the surroundings during the reaction.

ii. $n(C_4H_{10}) = \dfrac{m}{M} = \dfrac{450\text{ g}}{58.0\text{ g mol}^{-1}} = 7.75\text{ mol}$

2 mol C_4H_{10} produces 5 760 kJ

7.75 mol produces $\dfrac{5\,760\text{ kJ mol}^{-1}}{2} \times 7.75\text{ mol} = 2.23 \times 10^4\text{ kJ}$ (22 300 kJ)

d. $\Delta_r H = \sum E(\text{bonds broken}) - \sum E(\text{bonds made}) = -92\text{ kJ mol}^{-1}$

$\sum E(\text{bonds broken}) = E(N{\equiv}N) + 3 \times E(H{-}H)$

$= 945 + 3 \times 436$

$= 945 + 1\,308 = 2\,253\text{ kJ}$

$\sum E(\text{bonds made}) = 3 \times E(N{-}H)$

Substituting in $\Delta_r H = \sum E(\text{bonds broken}) - \sum E(\text{bonds made})$ gives:

$-92\text{ kJ mol}^{-1} = 2\,253\text{ kJ mol}^{-1} - 3 \times E(N{-}H)$

Rearranging gives:

$3 \times E(N{-}H) = 2\,253 + 92 = 2\,345\text{ kJ mol}^{-1}$

$E(N{-}H) = 782\text{ kJ mol}^{-1}$

e. $n(CH_3OH) = \dfrac{4\,428\text{ kJ}}{91\text{ kJ mol}^{-1}} = 48.7\text{ mol}$

$m(CH_3OH) = n \times M = 48.66 \times 32\text{ g mol}^{-1} = 1\,557\text{ g} = 1.56\text{ kg}$

If the mass of 1.00 L of methanol is 0.790 kg, then the volume of 1.56 kg is:

$V = \dfrac{1.56\text{ kg}}{0.790\text{ kg L}^{-1}} = 1.97\text{ L}$

(**A** – recognises exothermic/endothermic reactions with reasons, one step of calculations correct / energy diagram correct shape; **M** – explains endothermic and exothermic reactions, links endothermic reactions to breaking forces of attraction between molecules, method for energy calculations correct, energy diagram all correct; **E** – all calculations correct)

Question Two

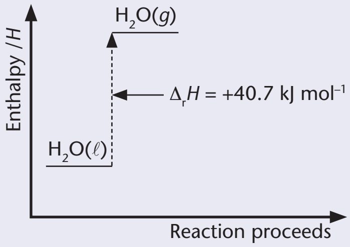
p. 30

a. **exothermic**

Temperature increase indicates that energy is released during the reaction.

b. i. Energy is required to change the state of the water from liquid to gas. The energy is used to break the forces of attraction between the molecules allowing them to separate, changing the substance from liquid to gas. Since the body heats up during exercise, more water can be evaporated.

ii.



c. $n(Fe_2O_3) = \dfrac{m}{M} = \dfrac{50.0\text{ g}}{160\text{ g mol}^{-1}} = 0.313\text{ mol}$

$n(CuO) = \dfrac{m}{M} = \dfrac{50.0\text{ g}}{79.6\text{ g mol}^{-1}} = 0.628\text{ mol}$

For Fe_2O_3: 1 mol releases 852 kJ

0.313 mol releases 0.313 × 852 = 266 kJ

For CuO: 1 mol releases $\frac{1}{3} \times 1\,520 = 507$ kJ

0.628 mol releases 0.628 × 507 = 318 kJ

Hence, 50.0 g of CuO releases more energy than 50.0 g of Fe_2O_3

d. **Reaction 1**

$\sum E(\text{bonds broken})$: $E(N{-}N) + 4 \times E(N{-}H) + E(O{=}O)$

$= 158 + 4 \times 391 + 498 = 2\,220\text{ kJ}$

$\sum E(\text{bonds made})$: $E(N{\equiv}N) + 4 \times E(O{-}H)$

$= 945 + 4 \times 463 = 2\,797\text{ kJ}$

$\Delta_r H = \sum E(\text{bonds broken}) - \sum E(\text{bonds made})$

$= 2\,220 - 2\,797 = -577\text{ kJ mol}^{-1}$

Reaction 2

$\sum E(\text{bonds broken})$: $E(N{-}N) + 4 \times E(N{-}H) + 2 \times E(F{-}F)$

$= 158 + 4 \times 391 + 2 \times 159 = 2\,040\text{ kJ}$

$\sum E(\text{bonds made})$: $E(N{\equiv}N) + 4 \times E(H{-}F)$

$= 945 + 4 \times 567 = 3\,213\text{ kJ}$

$\Delta_r H = \sum E(\text{bonds broken}) - \sum E(\text{bonds made})$

$= 2\,040 - 3\,213 = -1\,173\text{ kJ mol}^{-1}$

Hence, **Reaction 2** releases more energy.

(**A** – recognises exothermic/endothermic reactions with reasons, one step of calculations correct / energy diagram correct shape; **M** – explains endothermic and exothermic reactions, links endothermic reactions to breaking forces of attraction between molecules, method for energy calculations correct, energy diagram all correct; **E** – all calculations correct)

Question Three

p. 32

a. **endothermic**

Temperature decrease *or* heat/energy is absorbed from the surroundings.

b. **exothermic**

ΔH is negative *or* energy is released when the reaction occurs *or* energy of products is greater than energy of reactants.

c. i. Energy is needed to change the liquid to a gas. The energy is needed to break the forces of attraction between the pentane molecules.