CONTENTS

HOMEWORK

DATE	PARTICULARS

DATE	PARTICULARS

NOTES

NOTES

NOTES

NOTES

NOTES

3 Chemical Bonding

This topic introduces the different ways by which chemical bonding occurs and the effect this can have on physical properties.

3.1 Ionic Bonding ☐

3.4 Metallic Bonding ☐

3.5 Bonding and physical properties ☐

BONDING: OTHER

3 Chemical bonding

This topic introduces the different ways by which chemical bonding occurs and the effect this can have on physical properties.

	Learning outcomes Candidates should be able to:
3.1 Ionic bonding	a) describe ionic bonding, as in sodium chloride, magnesium oxide and calcium fluoride, including the use of 'dot-and-cross' diagrams
3.4 Metallic bonding	a) describe metallic bonding in terms of a lattice of positive ions surrounded by delocalised electrons
3.5 Bonding and physical properties	a) describe, interpret and predict the effect of different types of bonding (ionic bonding, covalent bonding, hydrogen bonding, other intermolecular interactions, metallic bonding) on the physical properties of substances b) deduce the type of bonding present from given information c) show understanding of chemical reactions in terms of energy transfers associated with the breaking and making of chemical bonds

IONIC BONDING

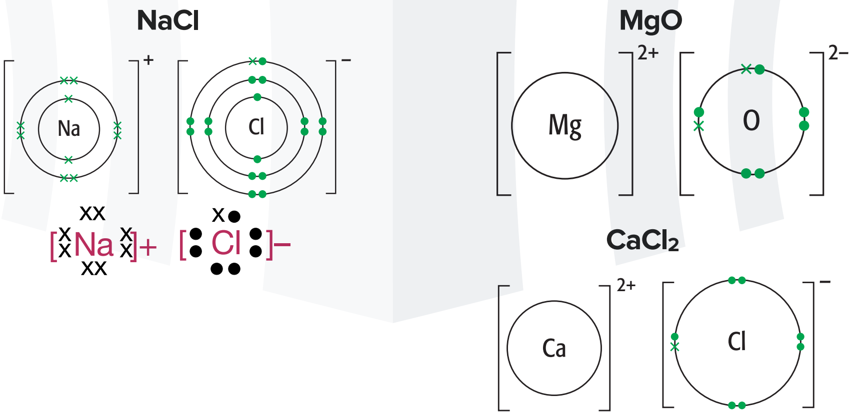
Ionic bonding involves a transfer of one or more electrons from one atom to another, leading to the formation of an ionic bond.

The strong electrostatic attraction that prevails between the oppositely charged ions in a crystal lattice is referred to as ionic bonding.

Positive ions, known as cations, are formed when electrons are removed from atoms. They are smaller than the original atom. The energy associated with the process is known as the ionisation energy

Negative ions, known as anions, are larger than the original atom. Energy is released as the nucleus pulls in an electron. This energy is the **electron affinity**.

‘DOT – AND – CROSS’ DIAGRAMS



SKILL CHECK

(a) Define ionic bonding.

(b) Explain in terms of electrons, how an **ionic bond** forms between atoms of **calcium** and atoms of **fluorine**.

(c) Draw electron configuration diagrams for a **calcium ion** and for a **fluoride ion**, showing their charges and outer electrons.

SKILL CHECK

Draw 'dot-and-cross' diagrams for:

(a) lithium fluoride

(b) magnesium chloride

(c) lithium oxide

(d) calcium oxide

SKILL CHECK

In a historically famous experiment Wohler heated “inorganic” ammonium cyanate in the absence of air. The only product of the reaction was “organic” urea, $CO(NH_2)_2$. No other products were formed in the reaction.

What is the formula of the cyanate ion present in ammonium cyanate?

A CNO^- **B** CNO^{2-} **C** CO^- **D** NO^-

IONIC COMPOUNDS

The compounds formed by ionic bonds do not contain individual molecules, but are formed of an infinite assembly of ions.

The ions due to their mutual attraction arrange themselves in a regular pattern. Thus they are always crystalline solids at room temperature.

The electrical force binding them being very strong, they are non–volatile with high melting and boiling points. Every bond in the lattice needs to be broken down to melt the ionic compound.

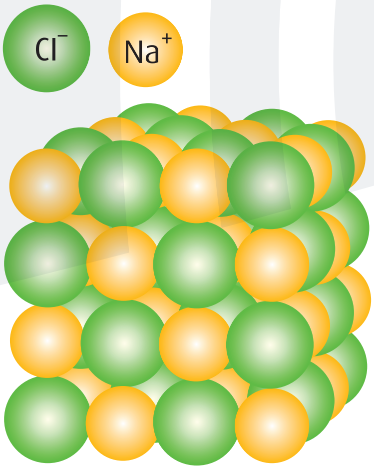
IONIC COMPOUNDS

Because they comprised of ions they conduct electricity in the molten or aqueous state, once the ions are made mobile.

In the solid state they do not conduct electricity. Generally ionic compounds are soluble in polar solvents like water.

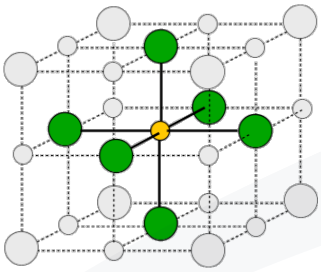
IONIC LATTICE

Oppositely charged ions held in a regular 3-dimensional lattice by electrostatic attraction

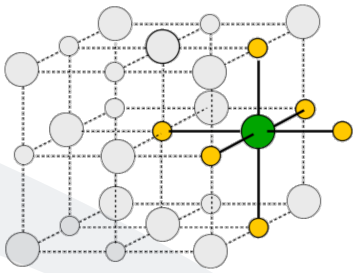


IONIC LATTICE

Each Na^+ is surrounded by 6 Cl^-



Each Cl^- is surrounded by 6 Na^+

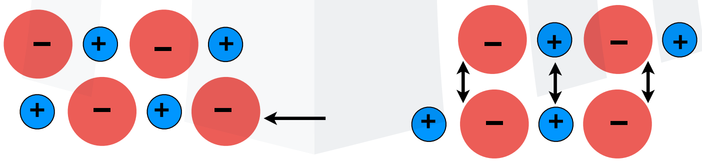


● Na^+ Sodium ion

● Cl^- Chloride ion

BRITTLE IONIC LATTICES

Ionic compounds are hard. However they are brittle. With a slight deforming force it is possible to slightly displace one layer of ions relative to the next and thereby bring ions of similar charge next to each other. Similar ions repel each other forcing apart the two portions of the crystal.



If you move a layer of ions, you get ions of the same charge next to each other. The laters repel each other and the crystal breaks up

SKILL CHECK

Magnesium oxide may be used for the lining of an electric furnace for making crockery. **Which properties of magnesium oxide help to explain this use?**

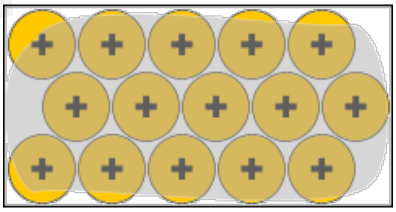
	strong forces between particles	ionic bonding	electrical conductor
A	yes	yes	no
B	yes	no	yes
C	no	yes	no
D	no	no	yes

METALLIC BONDING

Involves a lattice of positive ions surrounded by delocalised electrons

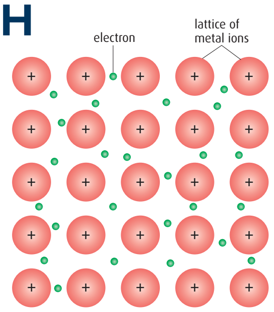
Metal atoms achieve stability by “off-loading” outer shell electrons to attain the electronic structure of the nearest noble gas. These electrons join up to form a mobile cloud which prevents the newly-formed positive ions from flying apart due to repulsion between similar charges.

Metals are excellent conductors of electricity because the ELECTRON CLOUD IS MOBILE, electrons are free to move throughout its structure. Electrons attracted to the positive end are replaced by those entering from the negative end.

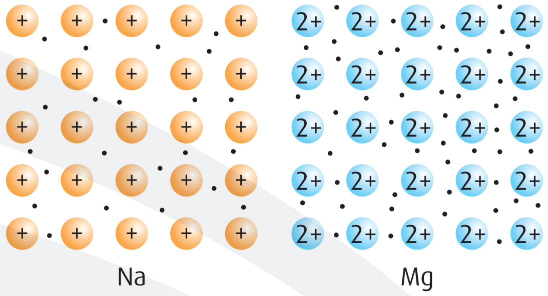


METALLIC BOND STRENGTH

The strength of the metallic bonding in sodium is relatively weak because each atom donates one electron to the cloud. The metallic bonding in potassium is weaker than in sodium because the resulting ion is larger and the electron cloud has a bigger volume to cover so is less effective at holding the ions together.

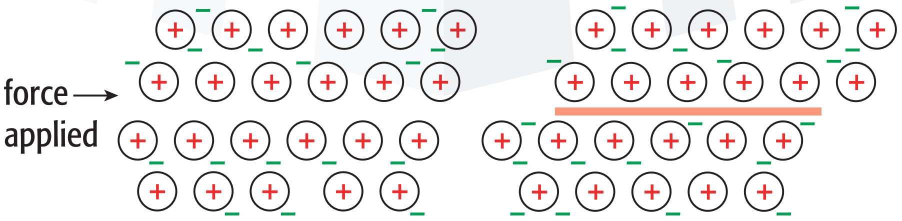


The metallic bonding in magnesium is stronger than in sodium because each atom has donated two electrons to the cloud. The greater the electron density holds the ions together more strongly.

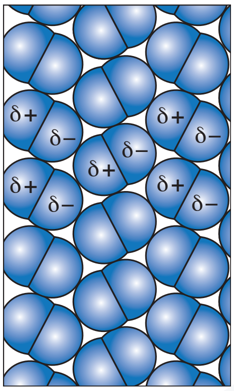


METALLIC BONDING

When force is applied, layers can slide over one another, since attractive forces between metal ions and sea of electrons in every direction, new metallic bonds are easily re-formed, attaining a different shape. This makes metals malleable and ductile.



SIMPLE MOLECULAR LATTICE



solid iodine – strong instantaneous dipole forces

Iodine also forms crystals with weak van der Waals’ forces between molecules.

This lattice is easily broken down when iodine is heated.

GIANT (MACRO) MOLECULES

Many atoms joined together in a regular array by a large number of covalent bonds, e.g. diamond, graphite, silicon (iv) oxide.

MELTING POINT	‣ Very high ‣ structure is made up of a large number of covalent bonds, all of which need to be broken if atoms are to be separated
ELECTRICAL CONDUCTIVITY	‣ Don’t conduct electricity - have no mobile ions or electrons ‣ but… Graphite conducts electricity
STRENGTH	‣ Hard - exists in a rigid tetrahedral structure, Diamond and silica (SiO2)… ‣ but Graphite is soft

DIAMOND

MELTING POINT

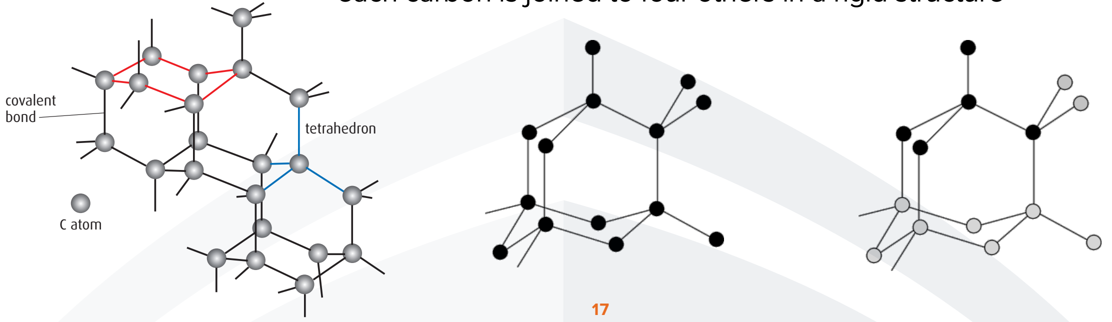
- VERY HIGH
- many covalent bonds must be broken to separate atoms

ELECTRICAL CONDUCTIVITY

- NON-CONDUCTOR
- No free electrons - all 4 carbon electrons used for bonding

STRENGTH

- STRONG
- each carbon is joined to four others in a rigid structure



SILICA

MELTING POINT

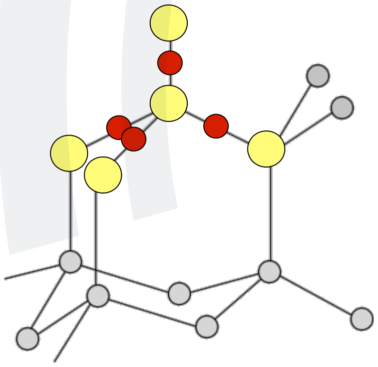
- VERY HIGH
- many covalent bonds must be broken to separate atoms

ELECTRICAL CONDUCTIVITY

- NON-CONDUCTOR
- No free electrons

STRENGTH

- STRONG
- each silicon atom is joined to four oxygens
- each oxygen atom are joined to two silicons



GRAPHITE

MELTING POINT

- VERY HIGH
- many covalent bonds must be broken to separate atoms

ELECTRICAL CONDUCTIVITY

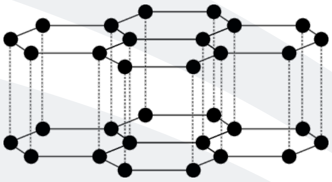
- CONDUCTOR
- Only three carbon electrons are used for bonding which leaves the fourth to move freely along layers

STRENGTH

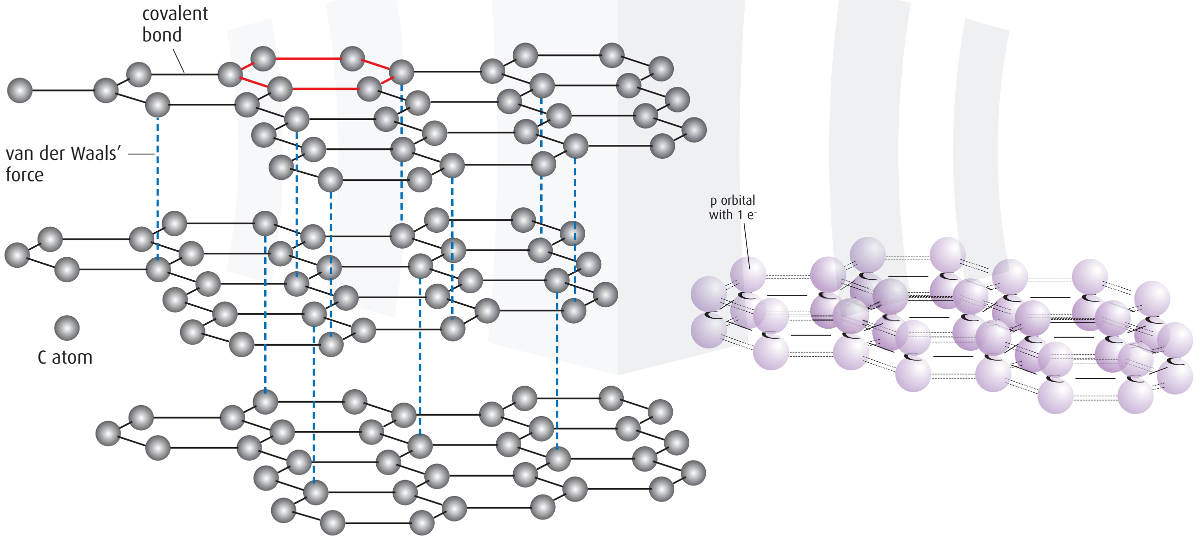
- SOFT
- each carbon is joined to three others in a layered structure
- layers are held by weak van der Waals' forces
- can slide over each other

Layers can slide over each other

Used as a lubricant and in pencils



GRAPHITE



CARBON NANOPARTICLES

Allotropes of carbon, known as fullerenes, have been recently discovered and possess unique and exciting properties.

They are based on rings of carbon, in hexagonal arrangement, similar to the structure of graphite.

They have dimensions between 0.1 to 100 nanometers.

FULLERENES

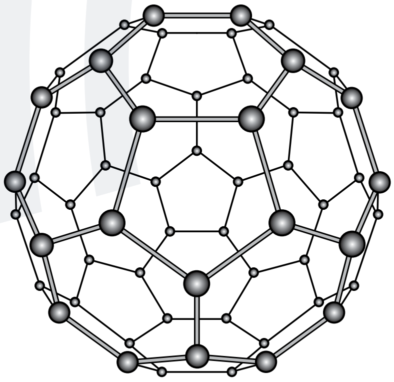
The firsts fullerene discovered was the buckminsterfullerene, C_{60}, which has the shape of a football.

Properties:

Relatively low sublimation point (weak van der Waals between each molecule).

Poor conductor of electricity (extent of electron delocalisation is lower).

More reactive compared to graphite. Reacts with hydrogen, fluorine, chlorine, bromine and oxygen.

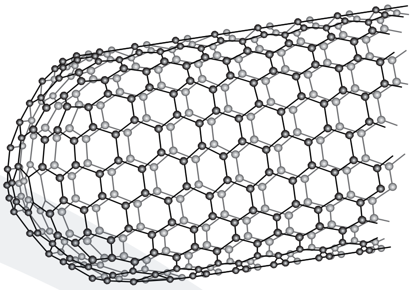


Buckminsterfullerene

NANOTUBES

Nanotubes are fullerenes hexagonally arranged carbon atoms like a single layer of graphite bent into a cylinder.

Properties:



High electrical conductivity along axis of cylinder.

High tensile strength.

Very high melting points (3500°C).

Used in tiny electrical circuits as wires, as electrodes in paper thin batteries, treatment of certain types of cancer in drug delivery and to strengthen clothing.

GRAPHENE

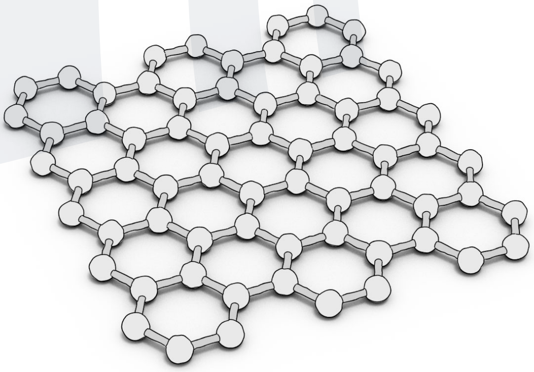
Single isolated layer of graphite.

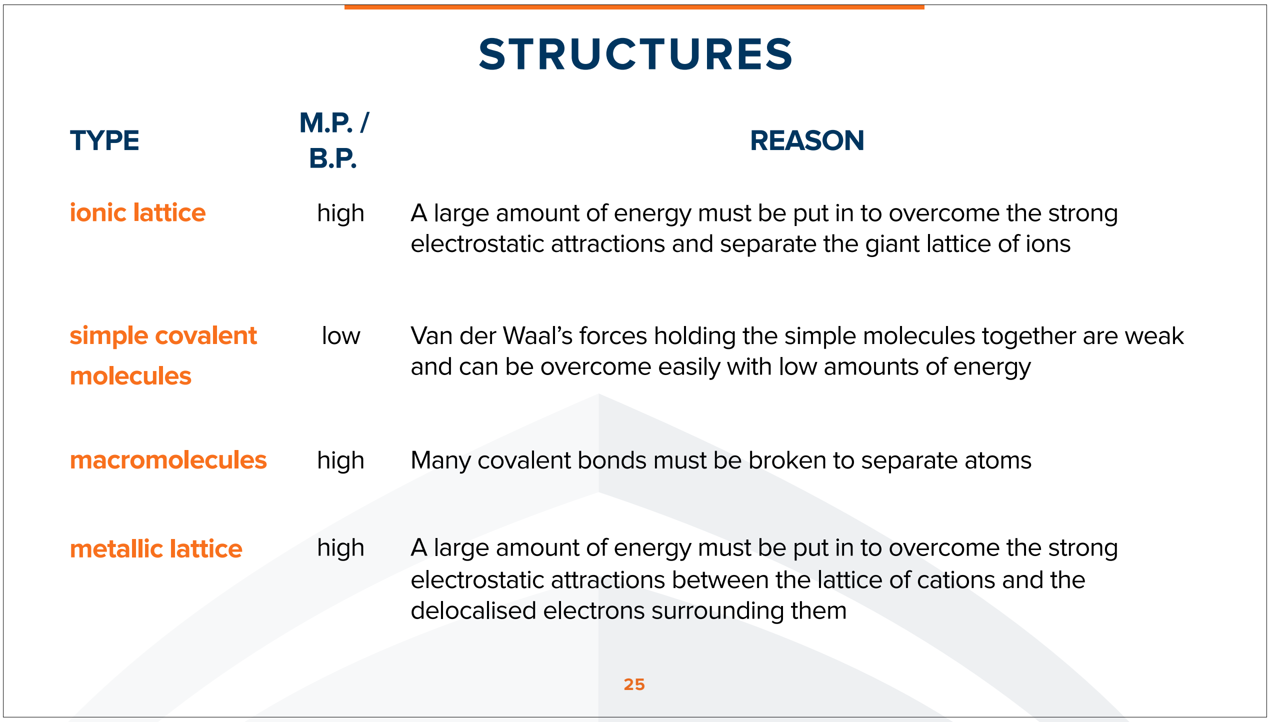
Not completely rigid and shape can be distorted.

Most reactive form of carbon (low melting point).

Extremely strong.

Conducts electricity.





STRUCTURES

TYPE	M.P. / B.P.	REASON
ionic lattice	high	A large amount of energy must be put in to overcome the strong electrostatic attractions and separate the giant lattice of ions
simple covalent molecules	low	Van der Waal's forces holding the simple molecules together are weak and can be overcome easily with low amounts of energy
macromolecules	high	Many covalent bonds must be broken to separate atoms
metallic lattice	high	A large amount of energy must be put in to overcome the strong electrostatic attractions between the lattice of cations and the delocalised electrons surrounding them

NOTES

OVERALL BONDING WS 1

SECTION A

1 Which of the following exists in the solid state as a giant covalent lattice?

A ice

B iodine

C silicon(IV) oxide

D tin(IV) chloride

2 Why does copper wire conduct electricity when a potential difference is applied?

A Bonding electrons in the crystal lattice move.

B Copper(II) ions move to the cathode.

C The atoms of copper become ionised.

D The crystal lattice breaks down.

3 When heated, solid iodine readily forms iodine vapour.

What does this information suggest about the nature of the particles in these two physical states of iodine?

	solid	*vapour*
A	ionic	atomic
B	ionic	molecular
C	molecular	atomic
D	molecular	molecular

4 Which of the following solids has a simple molecular lattice?

A magnesium oxide

B sodium

C silicon(IV) oxide

D sulphur

5 Why does the exothermic reaction

$$\text{C(diamond)} \longrightarrow \text{C(graphite)} \qquad \Delta H = -3\,\text{kJ mol}^{-1}$$

not occur spontaneously?

A A tetrahedral configuration is always more stable than a planar one.

B Diamond has only strong covalent bonds whereas graphite has both covalent bonds and van der Waals' forces.

C The change from diamond to graphite has a high activation energy.

D The density of graphite is less than that of diamond.

6 A substance commonly found in the house or garden has the following properties.

- It is combustible.
- It is an electrical insulator.
- It melts over a range of temperature.

What could the substance be?

A brass

B paper

C poly(ethene)

D silicon(IV) oxide

7 Which solid exhibits more than one kind of chemical bonding?

A brass

B copper

C diamond

D ice

8 Which type of bonding is **never** found in elements?

A covalent

B ionic

C metallic

D van der Waals' forces

[W'15 2 Q1]

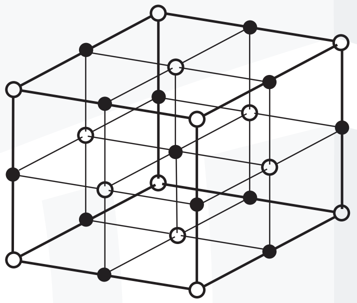
9 Three substances, *R*, *S*, *T*, have physical properties as shown.

substance	mp / °C	bp / °C	electrical conductivity	
			of solid	of liquid
R	801	1413	poor	good
S	2852	3600	poor	good
T	3550	4827	good	not known

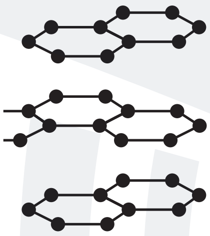
What could be the identities of *R*, *S* and *T* ?

	R	*S*	*T*
A	NaF	KC*l*	Cu
B	NaBr	BaO	SiO_2
C	NaC*l*	MgO	C [graphite]
D	NaBr	CaO	C [diamond]

10 The diagram shows part of the lattice structures of solids **X** and **Y**. [In **X**, ○ and ● represent particles of different elements.]



X



Y

What are the types of bonding present in **X** and **Y**?

	X	**Y**
A	covalent	metallic
B	ionic	covalent
C	ionic	metallic
D	metallic	ionic

11 What are the lattice structures of solid diamond, iodine and silicon(IV) oxide?

	giant molecular	simple molecular
A	diamond, silicon(IV) oxide	iodine
B	diamond, iodine	silicon(IV) oxide
C	iodine	diamond, silicon(IV) oxide
D	silicon(IV) oxide	diamond, iodine

12 Which solid has a simple molecular lattice?

A calcium fluoride

B nickel

C silicon(IV) oxide

D sulfur

13 Three substances, R, S and T, have physical properties as shown.

substance	R	S	T
mp / °C	801	2852	3550
bp / °C	1413	3600	4827
electrical conductivity of solid	poor	poor	good

What could be the identities of R, S and T?

	R	S	T
A	MgO	NaCl	C [graphite]
B	MgO	NaCl	SiO_2
C	NaCl	MgO	C [graphite]
D	NaCl	MgO	SiO_2

14 Three elements, **X**, **Y** and **Z**, have the physical properties shown in the table.

element	melting point / °C	boiling point / °C	density / g cm^{-3}
X	–7	59	3.12
Y	98	883	0.97
Z	649	1107	1.74

What could be the identities of **X**, **Y** and **Z**?

	X	**Y**	**Z**
A	Br_2	Al	Si
B	Br_2	Na	Mg
C	I_2	Mg	Na
D	I_2	Si	K

15 Three compounds have the physical properties shown in the table.

compound	**P**	**Q**	**R**
melting point / °C	2852	993	–119
boiling point / °C	3600	1695	39
conductivity (solid)	poor	poor	poor
conductivity (liquid)	good	good	poor
conductivity (aqueous)	insoluble	good	insoluble

What might be the identities of **P**, **Q** and **R**?

	P	**Q**	**R**
A	MgO	KCl	NH_3
B	MgO	NaF	C_2H_5Br
C	SiO_2	KCl	C_2H_5Br
D	SiO_2	NaF	HCl

16 Which solid contains more than one kind of bonding?

A iodine

B silicon dioxide

C sodium chloride

D zinc

17 Some car paints contain small flakes of silica, SiO_2.

In the structure of solid SiO_2

- each silicon atom is bonded to **x** oxygen atoms,
- each oxygen atom is bonded to **y** silicon atoms,
- each bond is a **z** type bond.

What is the correct combination of **x**, **y** and **z** in this statement?

	x	**y**	**z**
A	2	1	covalent
B	2	1	ionic
C	4	2	covalent
D	4	2	ionic

18 Which pair of elements has chemical bonds of the same type between their atoms in the solid state?

A aluminium and phosphorus

B chlorine and argon

C magnesium and silicon

D sulfur and chlorine

19 Two conversions are outlined below.

NH_4^+ NH_3

C_2H_4 C_2H_6

What similar feature do these two conversions have?

A a lone pair of electrons in the product

B change in oxidation state of an element

C decrease in bond angle of the species involved

D disappearance of a π bond

20 Magnesium oxide is used to line industrial furnaces because it has a very high melting point.

Which type of bond needs to be broken for magnesium oxide to melt?

A co-ordinate

B covalent

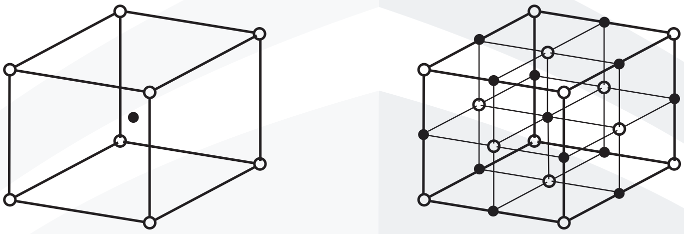
C ionic

D metallic

21 The table gives the radii, in pm, of some ions. [1 pm = 10^{-12} m]

ion	radii
Na^+	102
Mg^{2+}	72
Cs^+	167
Cl^-	181
O^{2-}	140

Caesium chloride, CsC*l*, has a different lattice structure from both sodium chloride, NaC*l*, and magnesium oxide, MgO.



CsC*l* lattice NaC*l* and MgO lattice

Which factor appears to determine the type of lattice for these three compounds?

A the charge on the cation

B the ratio of the ionic charges

C the ratio of the ionic radii

D the sum of the ionic charges

22 Substances X, Y and Z are all solids. Some of their physical properties are given in the table.

substance	X	Y	Z
melting point / °C	772	114	1610
boiling point / °C	1407	183	2205
electrical conductivity of the liquid state	conducts	does not conduct	does not conduct

What type of lattice could each substance have?

	X	Y	Z
A	giant molecular	simple molecular	ionic
B	ionic	giant molecular	simple molecular
C	ionic	simple molecular	giant molecular
D	simple molecular	ionic	giant molecular

23 Four substances have the physical properties shown.

Which substance is an ionic solid?

	melting point / °C	boiling point / °C	electrical conductivity of solid	electrical conductivity of molten substance	electrical conductivity of aqueous solution
A	–115	–85	poor	poor	good
B	660	2470	good	good	insoluble
C	993	1695	poor	good	good
D	1610	2230	poor	poor	insoluble

24 Which solid contains more than one kind of bonding?

A copper

B diamond

C ice

D magnesium oxide

25 $AlCl_3$ reacts with $LiAlH_4$ and $(CH_3)_3N$ to give $(CH_3)_3NAlH_3$.

Which statement about $(CH_3)_3NAlH_3$ is correct?

A It contains hydrogen bonding.

B It is dimeric.

C The Al atom has an incomplete octet of electrons.

D The bonds around the Al atom are tetrahedrally arranged.

26 Solid carbon dioxide, CO_2, is similar to solid iodine, I_2, in its structure and properties. Carbon is in Group 14. Silica, SiO_2, is a Group 14 compound.

Which statement about solid CO_2 and solid SiO_2 is correct?

A Both solids exist in a lattice structure.

B Both solids have a simple molecular structure.

C Both solids have atoms joined by single covalent bonds.

D Both solids change spontaneously to gas at s.t.p.

27 Which solid has a simple molecular lattice?

A calcium fluoride

B nickel

C silicon(IV) oxide

D sulfur

28 Some car paints contain small flakes of silica, SiO_2.

In the structure of solid SiO_2

- each silicon atom is bonded to **x** oxygen atoms,
- each oxygen atom is bonded to **y** silicon atoms,
- each bond is a **z** type bond.

What is the correct combination of **x**, **y** and **z** in these statements?

	x	y	z
A	2	1	covalent
B	2	1	ionic
C	4	2	covalent
D	4	2	ionic

29 Three substances have the physical properties shown in the table.

substance	melting point / °C	boiling point / °C	conductivity (solid)	conductivity (liquid)	conductivity (aqueous)
U	420	907	good	good	insoluble
V	993	1695	poor	good	good
W	–70	58	poor	poor	hydrolyses, resulting solution conducts well

What could be the identities of **U**, **V** and **W**?

	U	**V**	**W**
A	Na	KCl	$SiCl_4$
B	Na	NaF	C_2H_5Br
C	Zn	KCl	HCl
D	Zn	NaF	$SiCl_4$

30 Which solid contains more than one type of bonding?

A iodine

B silicon dioxide

C sodium chloride

D zinc

[S'18 1 Q6]

31 Which element shows the greatest tendency to form covalent compounds?

A boron

B magnesium

C neon

D potassium

[W'16 2 Q13]

32 Solid carbon dioxide, CO_2, is similar to solid iodine, I_2, in its structure and properties. Carbon is in Group 14. Silica, SiO_2, is a Group 14 compound.

Which statement about solid CO_2 and solid SiO_2 is correct?

A Both solids exist in a lattice structure.

B Both solids have a simple molecular structure.

C Both solids have atoms joined by single covalent bonds.

D Both solids change spontaneously to gas at s.t.p.

[M'16 Q8]

SECTION B

For each of the questions in this section, one or more of the three numbered statements **1** to **3** may be correct.

Decide whether each of the statements is or is not correct (you may find it helpful to put a tick against the statements that you consider to be correct).

The responses **A** to **D** should be selected on the basis of

A	B	C	D
1, 2 and **3** are correct	**1** and **2** only are correct	**2** and **3** only are correct	**1** only is correct

No other combination of statements is used as a correct response.

1 Which of the following are features of the structure of metallic copper?

1 ionic bonds

2 delocalised electrons

3 lattice of ions

2 Silicon tetrachloride, $SiCl_4$, is a liquid of low boiling point. In the presence of water it decomposes to form silicon(IV) oxide and hydrogen chloride.

What types of bonding occur in $SiCl_4$(l)?

1 co–ordinate bonding

2 covalent bonding

3 van der Waals forces

3 Which of the following statements are correct for the sequence of compounds below considered from left to right?

NaF MgO AlN SiC

1 The electronegativity difference between the elements in each compound increases.

2 The formula-units of these compounds are isoelectronic (have the same number of electrons).

3 The bonding becomes increasingly covalent.

4 Which pairs of compounds contain one that is giant ionic and one that is simple molecular?

1 Al_2O_3 and Al_2Cl_6

2 SiO_2 and $SiCl_4$

3 P_4O_{10} and PCl_3

5 Which of the following solids contain more than one type of chemical bond?

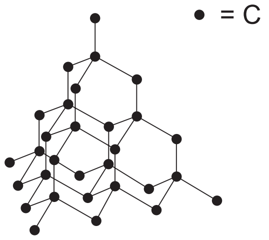
1 brass (an alloy of copper and zinc)

2 graphite

3 ice

6 Which diagrams represent part of a giant molecular structure?

1



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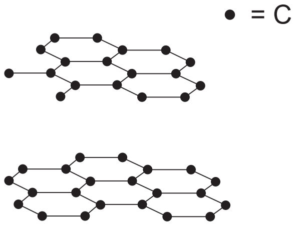

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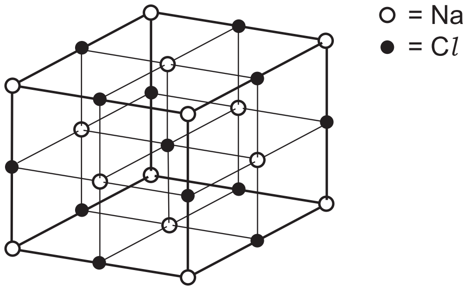


7 Boron is a non-metallic element which is placed above aluminium in Group III of the Periodic Table. It forms a compound with nitrogen known as boron nitride which has a graphite structure.

Which of the following conclusions can be drawn from this information?

1 The empirical formula of boron nitride is BN.

2 The boron and nitride atoms are likely to be arranged alternately in a hexagonal pattern.

3 Boron nitride has a layer structure with van der Waals' forces between the layers.

8 The Group IV elements carbon, silicon and germanium all exist in a diamond structure. The bond lengths in these structures are given below.

element *X*	C	Si	Ge
bond length *X*–*X* / nm	0.154	0.234	0.244

Why does the bond length increase down the group?

1 Orbital overlap decreases down the group.

2 Atomic radius increases down the group.

3 Nuclear charge increases down the group.

9 Which of these substances have a giant structure?

1 silicon(IV) oxide

2 baked clay found in crockery

3 phosphorus(V) oxide

10 Which substances have a giant structure?

1 calcium oxide

2 calcium

3 baked clay found in crockery

11 Which compounds contain covalent bonds?

1 aluminium chloride

2 ammonia

3 calcium fluoride

12 Which substances contain delocalised electrons?

1 cyclohexene

2 graphite

3 sodium

OVERALL BONDING WS 2

1 **(a)** Give the full electron configuration of the following.

(i) Mg ....

(ii) Mg^{2+} ....

(iii) O ....

(iv) O^{2-} ....

[2]

(b) **(i)** Describe, with the aid of a diagram, the lattice structure of magnesium oxide.

(ii) Use your diagram to interpret and explain **two** physical properties of magnesium oxide.

....

....

....

....

(iii) State **two** large scale uses of magnesium oxide.

....

....

[5]

2 **(a)** Salt, sodium chloride, forms transparent colourless crystals. Describe the bonding in sodium chloride crystals, give the formula of each particle and sketch part of the crystal structure.

[3]

(b) Explain why crystals of sodium chloride do not conduct electricity, but molten sodium chloride does.

....

....

....[2]

3 Drawing diagrams where appropriate, suggest in terms of structure and bonding, explanations for the following.

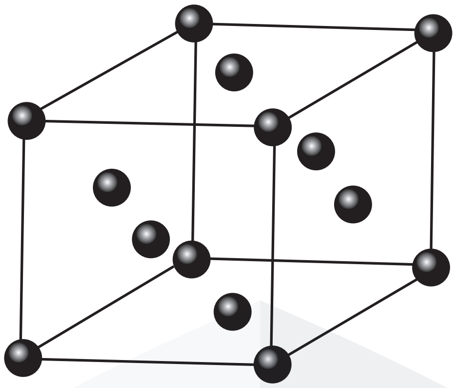
(i) the high melting point and boiling point of Al_2O_3

(ii) the low boiling point of SO_3

(iii) the melting point of SiO_2 is much higher than that of P_4O_{10}

[7]

4 Copper and iodine are both solids which have different physical and chemical properties. Each element has the same face-centred crystal structure which is shown below.



The particles present in such a crystal may be atoms, molecules, anions or cations. In the diagram above, the particles present are represented by ● .

(a) Which type of particles are present in the iodine crystal? Give their formula.

particle ....

formula .... [2]

(b) When separate samples of copper or iodine are heated to 50 °C, the copper remains as a solid while the iodine turns into a vapour.

(i) Explain, in terms of the forces present in the solid structure, why copper remains a solid at 50 °C.

....

....

....

(ii) Explain, in terms of the forces present in the solid structure, why iodine turns into a vapour when heated to 50°C.

....

....

....

[4]

5 Suggest, in terms of the structure and bonding, explanations for the following.
You should draw diagrams where you think they will help your answer.

(i) the high melting point of sodium chloride

(ii) the low melting point of silicon tetrachloride

[4]

6 The elements carbon and silicon are both in Group IV of the Periodic Table.
Carbon is the second most abundant element by mass in the human body and silicon is the second most common element in the Earth's crust.

Carbon and silicon each form an oxide of general formula XO_2.
At room temperature, CO_2 is a gas while SiO_2 is a solid with a high melting point.

(a) Briefly explain, in terms of the chemical bonds and intermolecular forces present in **each** compound, why CO_2 is a gas and SiO_2 is a solid at room temperature.

....

....

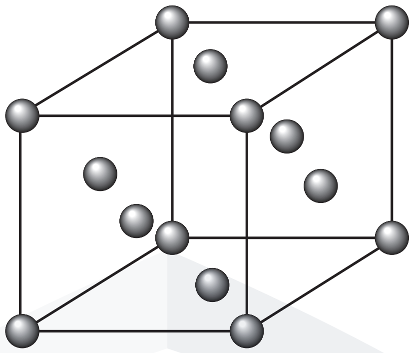
....

.... [3]

(b) Draw a simple diagram to show the structure of SiO_2. Your diagram should contain at least **two** silicon atoms **and** show clearly how many bonds each atom forms.

[2]

7 Copper, proton number 29, and argon, proton number 18, are elements which have different physical and chemical properties.
In the solid state, each element has the same face-centred cubic crystal structure which is shown below.



The particles present in such a crystal may be atoms, molecules, anions or cations. In the diagram above, the particles present are represented by ●.

(a) Which types of particle are present in the copper and argon crystals?
In each case, give their formula.

element	particle	formula
copper		
argon		

[2]

At room temperature, copper is a solid while argon is a gas.

(b) Explain these observations in terms of the forces present in **each** solid structure.

....

....

....

....

....

.... [4]

8 Separate samples of the oxides MgO and SiO_2 are melted.
Each molten sample is then tested to see whether or not it conducts electricity.

Suggest what would be the results in **each** case. Explain your answers.

MgO ....

....

....

SiO_2 ....

....

....

[4]

9 **(c)** Cerium is a lanthanoid metal that shows similar chemical reactions to some elements in the third period. Most of cerium's compounds contain Ce^{3+} or Ce^{4+} ions.

(i) Cerium shows the same structure and bonding as a typical metal.

Draw a labelled diagram to show the structure and bonding in cerium.

[2]

(ii) Cerium(IV) oxide, CeO_2, is a ceramic.

Suggest **two** physical properties of cerium(IV) oxide.

1. ....

2. ....

[2]

10 Sodium and silicon also react directly with chlorine to produce the chlorides shown.

chloride	melting point / °C	difference between the electronegativities of the elements
NaC*l*	801	2.2
SiCl_4	–69	1.3

(i) Describe what you would **see** during the reaction between sodium and chlorine.

....

....

.... [2]

(ii) Explain the differences between the melting points of these two chlorides in terms of their structure **and** bonding. You should refer to the difference between the electronegativities of the elements in your answer.

NaC*l* structure **and** bonding ....

....

SiCl_4 structure **and** bonding ....

....

explanation ....

....

....

....

....

.... [4]

11 Structure and bonding can be used to explain many of the properties of substances.

(a) Copper, ice, silicon(IV) oxide, iodine and sodium chloride are all crystalline solids.

Complete the table with:

- the name of a type of bonding found in each crystalline solid,
- the type of lattice structure for each crystalline solid.

crystalline solid	type of bonding	type of lattice structure
copper		
ice		
silicon(IV) oxide		
iodine		
sodium chloride		

[5]

(b) **(i)** Name the strongest type of intermolecular force in ice.

.... [1]

(ii) Draw a fully labelled diagram of two water molecules in ice, showing the force in **(i)** and how it forms.

[3]

5 Chemical energetics

Enthalpy changes and entropy changes accompany chemical reactions. This topic demonstrates why some reactions and processes are spontaneous and others are not.

5.1 Enthalpy change, Δ*H* ☐

5.2 Hess' Law, including Born-Haber cycles ☐

ENERGETICS

5 Chemical energetics

Enthalpy changes and entropy changes accompany chemical reactions. This topic demonstrates why some reactions and processes are spontaneous and others are not.

Learning outcomes
Candidates should be able to:

5.1 Enthalpy change, ΔH

a) explain that chemical reactions are accompanied by energy changes, principally in the form of heat energy; the energy changes can be exothermic (ΔH is negative) or endothermic (ΔH is positive)

b) explain and use the terms:
  - (i) *enthalpy change of reaction* and *standard condition*s, with particular reference to: formation, combustion, hydration, solution, neutralisation, atomisation
  - (ii) *bond energy* (ΔH positive, i.e. bond breaking)
  - **(iii) *lattice energy* (ΔH negative, i.e. gaseous ions to solid lattice)**

c) calculate enthalpy changes from appropriate experimental results, including the use of the relationship

$$\text{enthalpy change, } \Delta H = -mc\Delta T$$

d) explain, in qualitative terms, the effect of ionic charge and of ionic radius on the numerical magnitude of a lattice energy

5.2 Hess' Law, including Born-Haber cycles

a) apply Hess' Law to construct simple energy cycles, and carry out calculations involving such cycles and relevant energy terms, with particular reference to:
  - (i) determining enthalpy changes that cannot be found by direct experiment, e.g. an enthalpy change of formation from enthalpy changes of combustion
  - (ii) average bond energies
  - **(iii) the formation of a simple ionic solid and of its aqueous solution**
  - **(iv) Born-Haber cycles (including ionisation energy and electron affinity)**

b) construct and interpret a reaction pathway diagram, in terms of the enthalpy change of the reaction and of the activation energy

ENTHALPY CHANGES

Every chemical reaction involves the exchange of energy between (what we call) **the system** and **its surroundings.** For example, when magnesium and sulphuric acid react in a test-tube:

1. **The system** consists of the chemical bonds and other forces of attraction and repulsion between the atoms of reactants and of the products formed.
2. **The surroundings** include:
   a. The test tube.
   b. The air around the test tube.
   c. The thermometer dipped into the reacting mixture etc…

ENTHALPY CHANGES

The **total heat (energy) content** of any substance at constant pressure is called its **enthalpy**. Because of how difficult it is to observe things at the atomic and sub-atomic level, it is not possible to measure the **total enthalpy** of a substance.

A good analogy for this is trying to measure the volume of the Pacific Ocean. The structure of the seabed is so diverse it is practically impossible to measure just how much volume the Pacific Ocean has. **However, you can easily measure changes in the volume of the Ocean.**

Similarly, we can easily measure the **enthalpy change** that accompanies reactions. This is written as **ΔH** and pronounced ‘delta H.’ ΔH can either have a positive or a negative value, depending on the nature of energy exchange in the reaction.

ENTHALPY CHANGES

When energy flows from **the system** to **the surroundings**, the temperature of the surroundings **increases.** This indicates that heat was **given out** as a result of the reaction, and so the products of the reaction are at a lower energy level than the reactants were. These types of reactions are called **exothermic reactions.**

The value for ΔH is *negative* for exothermic reactions.

When energy flows from **the surroundings** to **the system**, the temperature of the surroundings **decreases.** This indicates that heat was **taken in** as a result of the reaction, and so the products of the reaction are at a higher energy level than the reactants were. These types of reactions are called **endothermic reactions.**

The value for ΔH is *positive* for endothermic reactions.

SKILL CHECK

Classify the following reactions as exothermic or endothermic:

1. Burning magnesium in excess oxygen.

2. Crystallisation of copper (II) sulphate from a saturated solution.

3. The thermal decomposition of magnesium nitrate.

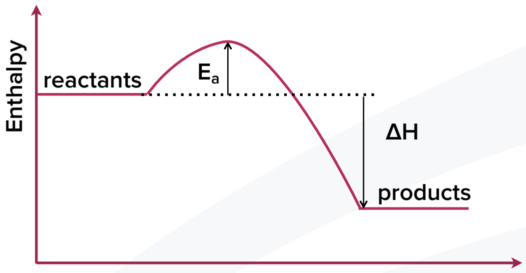
4. Fermentation of glucose by yeast.

5. Evaporation of sea water.

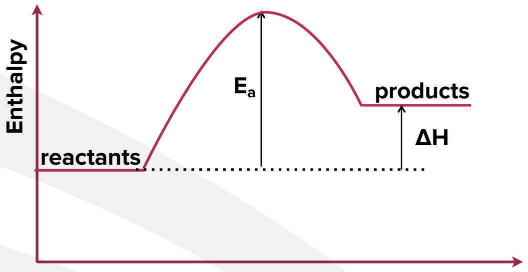
ENTHALPY CHANGES

To summarise:

Enthalpy change (ΔH) = Enthalpy of products – Enthalpy of reactants



Enthalpy of reactants > products,ΔH = —ve
EXOTHERMIC: Heat given out



Enthalpy of reactants < products, ΔH = +ve
ENDOTHERMIC: Heat absorbed

ENTHALPY CHANGES

We can draw **enthalpy profile diagrams** (also known as **reaction pathway diagrams**) to show enthalpy changes, as shown on the previous slide.

1. The enthalpy of the reactants and products is shown on the y-axis.

2. The x-axis shows the reaction pathway, with reactants on the left and products on the right.

SKILL CHECK

Draw enthalpy profile diagrams for:

a. Combustion of sulphur to form sulphur dioxide.

b. The neutralisation of potassium hydroxide by nitric acid.

STANDARD ENTHALPY CHANGES

The energy change that accompanies a reaction depends on:

1. The amount of reactants.
2. The physical state of the reactants and products.
3. The temperature of the reactants and products.

As reactants can exist at a variety of energy levels at different physical states, temperatures and pressures, there was a need to standardise conditions under which these values were measured. These are now known as **standard conditions.** They are a *pressure* of 10^5 Pa (100 KPa), *temperature* of 298K (25 °C) and each substance involved in the reaction is in its normal state at these values of temperature and pressure.

STANDARD ENTHALPY CHANGES

Enthalpy change measured under standard conditions is known as **standard enthalpy change.** The ⊖ symbol indicates that the enthalpy change refers to a reaction carried out under standard conditions. The unit of measurement for standard enthalpy change is the kilojoule per mole, written as **kJmol⁻¹.**

The equation below shows us that when one mole of methane gas reacts with two moles of oxygen gas to form one mole of carbon dioxide gas and two moles of water in the liquid state the **standard enthalpy change** is –890.3 kJmol-1.

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l) \qquad \Delta H^{\ominus} = -890.3\,\text{kJ mol}^{-1}$$

STANDARD ENTHALPY CHANGES

Enthalpy changes are described depending on what type of reaction is taking place. The following are some of the most commonly used enthalpy changes:

1. Standard enthalpy change of formation.
2. Standard enthalpy change of neutralisation.
3. Standard enthalpy change of combustion.
4. Standard enthalpy change of atomisation.

STANDARD ENTHALPY CHANGE OF FORMATION

Definition The enthalpy change when **one mole** of a compound is formed **in its standard state** from its **elements in their standard states** under **standard conditions.**

Symbol $\Delta H^{\ominus}_{f}$

Values Usually, but not exclusively, exothermic.

Examples

$$C_{(graphite)} + O_2(g) \longrightarrow CO_2(g)$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l)$$

$$2C_{(graphite)} + \tfrac{1}{2}O_2(g) + 3H_2(g) \longrightarrow C_2H_5OH(l)$$

Notes There should be only **one mole** of product on the right side of the equation.

Elements in their standard states have zero enthalpy of formation

Carbon is usually taken as the graphite allotrope

STANDARD ENTHALPY CHANGE OF COMBUSTION

Definition The enthalpy change when **one mole** of a substance undergoes **complete** combustion **under standard conditions**. All reactants and products are in their **standard states**.

Symbol $\Delta H^{\ominus}_{c}$

Values Always exothermic

Examples

$$C_{(graphite)} + O_2(g) \longrightarrow CO_2(g)$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \longrightarrow H_2O(l)$$

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$$

Notes There should only **one mole** of what is being combusted on the left side of the equation.

To aid balancing the equation, remember that you get **one carbon dioxide** molecule for **every carbon atom** in the original and **one water molecule** for every **two hydrogen atoms**. When you have done this, go back and balance the oxygen.

STANDARD ENTHALPY CHANGE OF NEUTRALISATION

Definition	The enthalpy change when an **acid** and a **base** react to form **one mole** of water under **standard conditions**.
Symbol	$\Delta H^{\ominus}_{neutralisation}$
Values	Exothermic
Example	$H^+ (aq) + OH^- (aq) \longrightarrow H_2O(l)$
Notes	A value of -57kJmol^{-1} is obtained when strong acids react with strong alkalis.

SKILL CHECK

Classify the following reactions as standard enthalpy changes of formation, combustion or neutralisation:

1. $MgCO_3 (s) \longrightarrow MgO (s) + CO_2 (g)$

2. $C_{(graphite)} + O_2 (g) \longrightarrow CO_2 (g)$

3. $HCl (g) + NH_3 (g) \longrightarrow NH_4Cl (s)$

4. $H_2 (g) + \frac{1}{2}O_2 (g) \longrightarrow H_2O (l)$

MEASURING ENTHALPY CHANGES

Calorimetry is used to measure enthalpy changes.

It usually involves heating (or cooling) known amounts of water.

Many reactions take place in aqueous solution – reactants and/or products dissolved in water - so water is the natural choice to measure a variety of enthalpy changes.

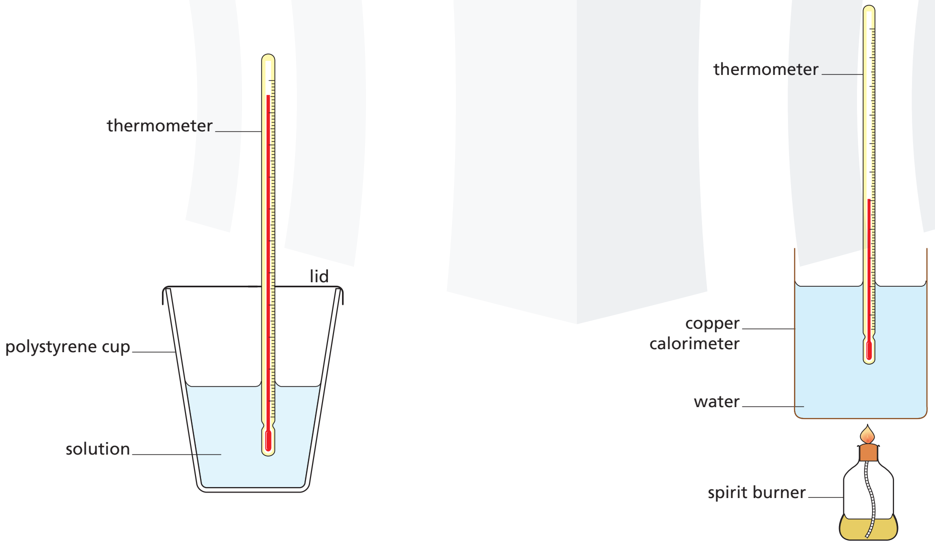
The enthalpy change of reaction causes the temperature of the water to change:

Endothermic reactions take in energy, temperature of water **falls**.

Exothermic reactions give energy out, temperature of water **rises**.

15

SIMPLE CALORIMETERS



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MEASURING ENTHALPY CHANGES

The energy required to change the temperature of a substance can be calculated using:

$$q = m \times c \times \Delta T$$

where **q** = heat energy, measured in Joules (J)

m = mass, measured in grams (g)

ΔT = change in temperature, measured in Kelvin (K) or degree Celsius (°C)

c = Specific Heat Capacity, measured in $JK^{-1}g^{-1}$ or $J°C^{-1}g^{-1}$

MEASURING ENTHALPY CHANGES

We can measure the temperature change using a thermometer marked in degrees Celsius, because a change in temperature is the same on either the Celsius or the Kelvin scale. The value for the specific heat capacity of water is 4.18 $JK^{-1}g^{-1}$ (or $J°C^{-1}g^{-1}$).

If we are dealing with aqueous solutions rather than pure water, it is the mass of water in the solution that should be included as the 'm' in this equation. Since the solutions are often quite dilute, and since the volume and heat capacity of a dilute solution are about the same as those of the water it contains, we normally use the volume of the solution rather than its mass to give this equation:

$$q = \text{volume} \times c \times \Delta T$$

CALCULATING ENTHALPY CHANGES

Step 1: work out the enthalpy change during the reaction

→ **Step 2**: work out how many moles of water are made from the reaction

→ **Step 3**: work out how much energy is evolved per mole of water

CALCULATING ENTHALPY CHANGES

25cm^3 of 1 moldm^{-3} HCl (aq) was neutralized by 25 cm^3 of 1 moldm^{-3} NaOH (aq). The temperature was increased by 5°C. Calculate the enthalpy change of neutralisation.

Enthalpy change $q = mc\,\Delta T$ *where* $m = 50$ g, $c = 4.2$ J°C^{-1}g^{-1}, $\Delta T = 5.0$°C

So, enthalpy change = $50 \times 4.2 \times 5.0 = 1050$ J

$$\underset{\textit{1 mole}}{HCl(aq)} + \underset{\textit{1 mole}}{NaOH(aq)} \longrightarrow NaCl(aq) + H_2O(l)$$

Number of moles of H_2O = number of moles of HCl = $0.025 \times 1.0 = 0.025$.

So, 0.025 moles of water are made.

CALCULATING ENTHALPY CHANGES

Number of moles of H_2O = number of moles of HCl = 0.025 x 1.0 = 0.025. So, 0.025 moles of water are made.

Enthalpy change when 0.025 moles of water made = 1050J

Enthalpy change when 1 mole of water made = 1050/0.025 = 42000 J

Enthalpy change of neutralisation = — 42 kJmol^{-1}

Heat energy gained by the water is lost by the reaction so change the sign!

CALCULATING ENTHALPY CHANGES

0.16g of methanol was burned in a spirit burner. 50g of water was heated from 19°C to 23°C.

(a) Calculate the enthalpy change of combustion of methane.

$q = mc\,\Delta T$ m = 50g, c = 4.2 $JK^{-1}g^{-1}$, $\Delta T = 4.0°C$

So, enthalpy change of water = 50 x 4.2 x 4.0 = 840 J

moles of methanol burned = mass/mass of 1 mole = 0.16/32 = 0.005 moles

CALCULATING ENTHALPY CHANGES

Enthalpy change of water for 1 mole of methane= 840/0.005 = 168000J = 168kJ

Enthalpy change of combustion = —168 kJ mol^{-1}

(b) Comment on why this experimental value is lower than the data booklet value.

Energy is lost to the surroundings, so that the temperature increase of the water is lower than that under ideal conditions.

CALCULATING ENTHALPY CHANGES

On complete combustion, 0.18g of hexane raised the temperature of 100g water from 22°C to 47°C. Calculate its enthalpy of combustion.

Heat absorbed by the water (q) = 100 x 4.18 x 25 = 10450 J = 10.45 kJ

Moles of hexane burned = mass/Mr = 0.18/86 = 0.00209

Enthalpy change = heat energy/moles = 10.45/0.00209 = 5000 kJ mol^{-1}

CALCULATING ENTHALPY CHANGES

25cm^3 of 2.0M HCl was added to 25cm^3 of 2.0M NaOH in an insulated beaker. The initial temperature of both solutions was 20°C. The highest temperature reached by the solution was 33°C. Calculate the Molar Enthalpy of Neutralisation.

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

Temperature rise (ΔT) = 306K — 293K = 13K

Volume of resting solution = 25 + 25 = 50cm^3 = 0.05dm^3

Equivalent mass of water = 50g (density is 1 gcm^{-3})

CALCULATING ENTHALPY CHANGES

Heat absorbed by the water (q) = m x c x ΔT = 500 x 4.18 x 13 = 2717 J = 2.717 kJ

Moles of HCl reacting = 2 x 25/1000 = 0.05 mol

Moles of NaOH reacting = 2 x 25/1000 = 0.05 mol

Moles of water produced = 0.05 mol

Enthalpy change per mole (ΔH) = heat energy/moles of water = 2.717/0.05 = 54.34 kJmol^{-1}

CALCULATING ENTHALPY CHANGES

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Moles of water produced = 0.05 mol

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SKILL CHECK

75cm^3 of 2.0 moldm^{-3} ethanoic acid, CH_3CO_2H, were placed in a plastic cup. The temperature was 18.2°C. To this were added 75cm^3 of 2.0 moldm^{-3} ammonium hydroxide, NH_4OH, whose temperature was 18.6 °C. After mixing, the highest temperature was 31.0 °C.

Calculate ΔH for the neutralisation.

29

SKILL CHECK

A burner containing hexanol, $C_6H_{13}OH$, had a mass of 325.68g. It was lit and placed under a copper calorimeter containing 250 cm^3 of water. The temperature of the water rose from 19.2°C to 31.6°C. Afterwards the burner's mass was 324.37 g.

Calculate:

A. The heat evolved.

B. ΔH_c for hexanol.

30

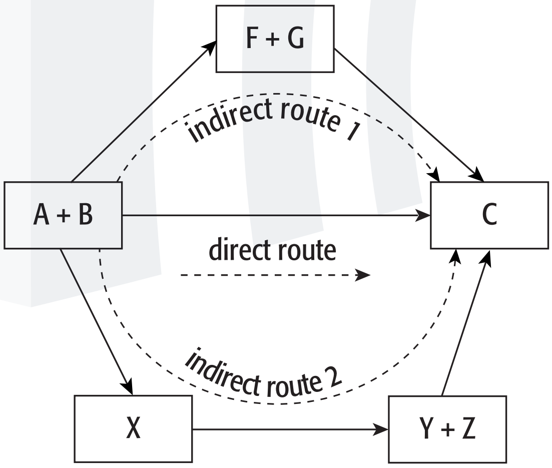
HESS’ LAW

There are very few reactions whose enthalpy change can be measured directly by measuring the change in temperature in a calorimeter. We can find enthalpy changes for other reactions indirectly. To do this we make use of **Hess’ Law**.

Hess’ law is an application of the first law of thermodynamics, which in its simplest form is the law of conservation of energy. Since energy cannot be destroyed or created the energy change in a chemical process should be the same as long as the initial and final states are the same.

HESS’ LAW

Hess’ law states that ‘the total enthalpy change in a chemical reaction is independent of the route by which the chemical reaction takes place as long as the initial and final conditions are the same’.



ΔH(direct) = ΔH(indirect 1) = ΔH(indirect 2)

HESS’ LAW

Some enthalpy changes are easy to measure experimentally, e.g.:

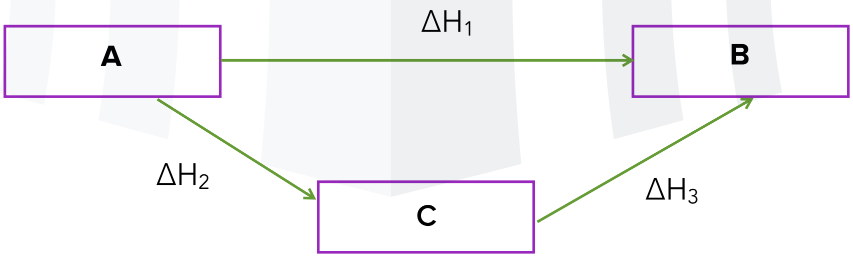
- ΔH°_c, enthalpy change of combustion.
- Some values for ΔH°_f, enthalpy change of formation.

These can be used to calculate some enthalpy changes that are difficult to measure experimentally e.g. :

- Some values for ΔH°_f, enthalpy change of formation.
- Some values for ΔH°_r, enthalpy change of reaction.

HESS’ LAW

This can be shown by making a Hess cycle as shown below:



$$A \rightarrow B \quad = \quad A \rightarrow C \quad + \quad C \rightarrow B$$

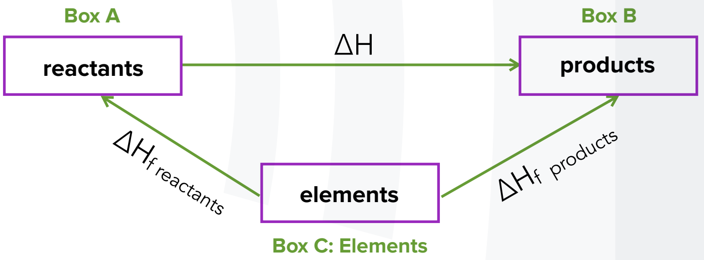
$$\Delta H_1 \quad = \quad \Delta H_2 \quad + \quad \Delta H_3$$

HESS’ LAW

1. Write out balanced equations for **all energy changes** mentioned in question.
2. Choose the most complicated (or populated) equation.
3. Fill the first two boxes with it.
4. Put everything else in the third box.
5. Balance each box to ensure the number of atoms in each box is the same.
6. The box with which your cycle begins is the start box and the one where the cycle ends is the stop box.
7. Construct an equation and solve.

Note: The enthalpies given are per mole therefore as the number of moles change, the energies have to be multiplied by the same number.

SUMMARY CYCLES



FOR FORMATION REACTIONS

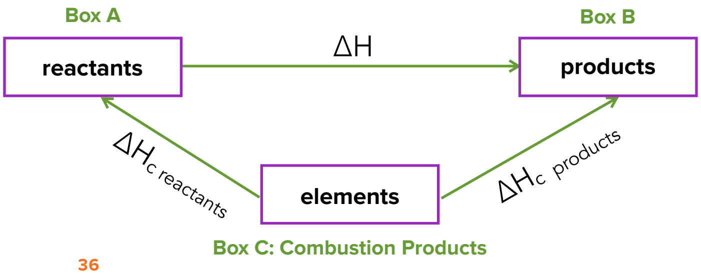
Start box: C **End box: B**

C ➜ A + A ➜ B = C ➜ B

FOR COMBUSTION REACTIONS

Start box: A **End box: C**

A ➜ B + B ➜ C = A ➜ C



EXAMPLE 1

Consider the following two reactions:

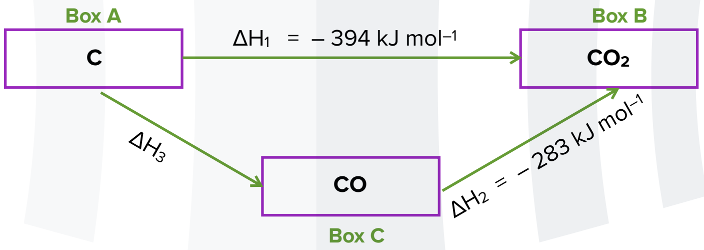
$C_{(graphite)} + O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_1 = -394\ \text{kJ mol}^{-1}$

$CO(g) + \frac{1}{2}O_2(g) \longrightarrow CO_2(g) \qquad \Delta H_2 = -283\ \text{kJ mol}^{-1}$

We can use Hess's Law to find the value of the incomplete combustion of carbon to form CO.

$C_{(graphite)} + \frac{1}{2}O_2(g) \longrightarrow CO(g) \qquad \Delta H_3 = ?$

EXAMPLE 1



Box A

C

$\Delta H_1 = -394\ \text{kJ mol}^{-1}$

Box B

CO_2

ΔH_3

CO

Box C

$\Delta H_2 = -283\ \text{kJ mol}^{-1}$

$A \rightarrow B = A \rightarrow C + C \rightarrow B$

$\Delta H_1 = \Delta H_3 + \Delta H_2$

$-394 = \Delta H_3 + (-283)$

$\Delta H_3 = ?$

EXAMPLE 2

Calculate the standard enthalpy of formation of methane from the following data:

$\Delta H^\circ_c(C) = -393\ kJ\ mol^{-1}$

$\Delta H^\circ_c(H2) = -286\ kJ\ mol^{-1}$

$\Delta H^\circ_c(CH4) = -890\ kJ\ mol^{-1}$

Step 1: form the equation of the formation of methane from its elements in their standard state.

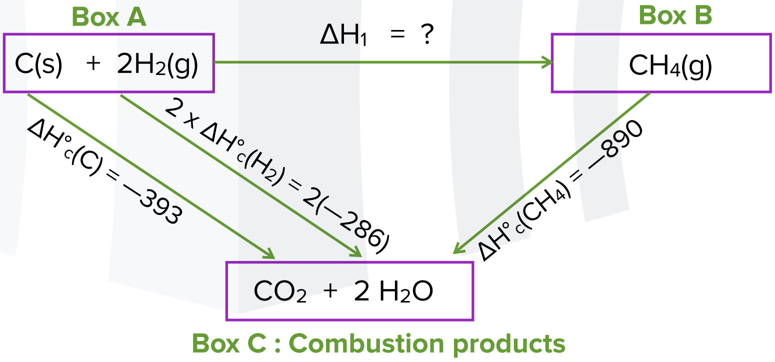
$$C_{(s)} + 2H_{2(g)} \longrightarrow CH_{4(g)}$$

EXAMPLE 2

Step 2: enter the reactants (in Box A) and products (in Box B) of Step 1 into the Hess cycle

Step 3: enter the combustion products of the Box A and Box B into Box C

Step 4: work out ΔH_1 from $\Delta H_1 = \Delta H_2 + \Delta H_3$



$A \rightarrow B + B \rightarrow C = A \rightarrow C$

$\Delta H_1 + \Delta H^\circ_c(CH_4) = \Delta H^\circ_c(C) + 2 \times \Delta H^\circ_c(H_2)$

$\Delta H_1 + (-890) = (-393) + 2(-286)$

$\Delta H_1 = ?$

SKILL CHECK

Calculate the standard enthalpy change for the following reaction, given that the standard enthalpies of formation of water, nitrogen dioxide and nitric acid are —286, +33 and —173 kJ mol^{-1} respectively. The value for oxygen is ZERO as it is an element (—252 kJ mol^{-1})

$$2H_2O(l) + 4NO_2(g) + O_2(g) \longrightarrow 4HNO_3(l)$$

SKILL CHECK

Calculate the standard enthalpy change of combustion of ethane from the following data:

- Standard enthalpy change of formation of carbon dioxide = – 394 kJ mol^{-1}
- Standard enthalpy change of formation of water = – 288 kJ mol^{-1}
- Standard enthalpy change of formation of ethane = – 92 kJ mol^{-1}

SKILL CHECK

Calculate the enthalpy change for the reaction:

$$4CH_3NHNH_2(l) + 5N_2O_4(l) \longrightarrow 4CO_2 + 12H_2O + 9N_2$$

Given the following data:

- Standard enthalpy change of formation of carbon dioxide = – 394 kJ mol^{-1}
- Standard enthalpy change of formation of water = – 288 kJ mol^{-1}
- Standard enthalpy change of formation of CH_3NHNH_2 = +53 kJ mol^{-1}
- Standard enthalpy change of formation of N_2O_4 = – 20 kJ mol^{-1}

BOND ENTHALPY

Bond Enthalpy is the energy required to break one mole of a particular type of bond in gaseous molecules under standard conditions.

Bond enthalpy values are all positive, showing that bond enthalpies are endothermic (breaking bonds takes in energy).

The values found in the data booklet are average values for each type of bond. The same bond may have very slightly different bond enthalpies in different compounds.

The O—H bond has a slightly different enthalpy value in each compound.

BOND ENTHALPY

When a particular bond is broken, energy is taken in (*endothermic*):

$$\text{H—H(g)} \longrightarrow \text{H(g)} + \text{H(g)} \qquad \mathbf{\Delta H = +436\ kJ\ mol^{-1}}$$

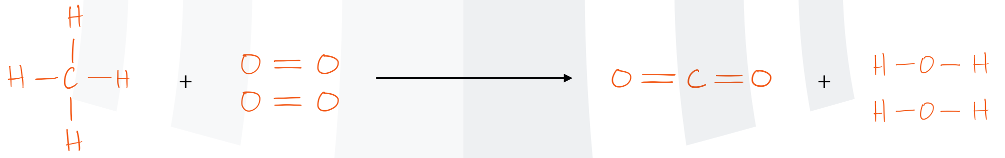
When the same bond is made, energy is given out (*exothermic*):

$$\text{H(g)} + \text{H(g)} \longrightarrow \text{H—H(g)} \qquad \mathbf{\Delta H = -436\ kJ\ mol^{-1}}$$

Note: The amount of energy needed to break or form the same type of bond is the same.

EXAMPLE

Calculate the enthalpy change of reaction when 1 mole of methane, CH_4 , burns completely to form carbon dioxide and water.



4 x C—H bonds

2 x O=O bonds

2 x C=O bonds

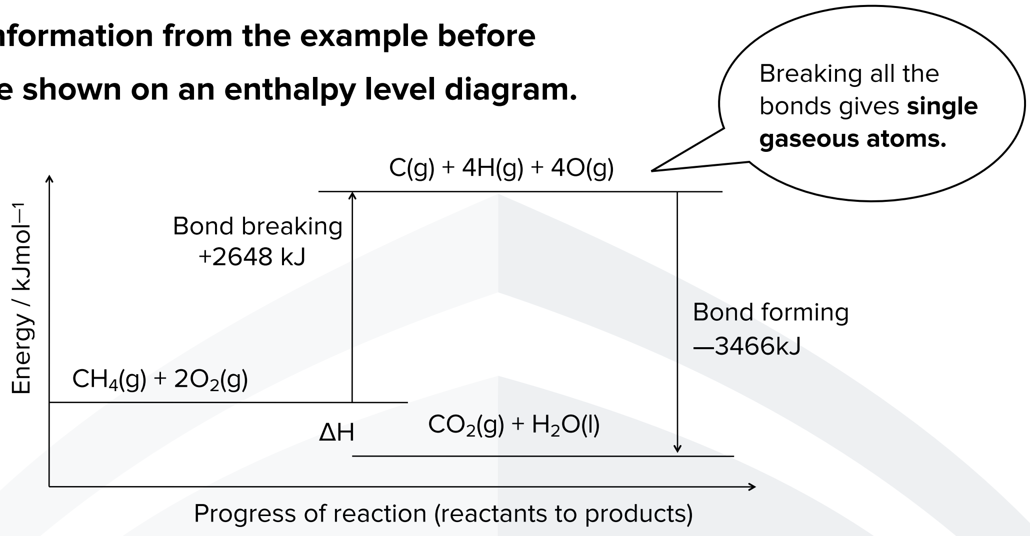
4 x O—H bonds

Reactant Bonds	Energy needed to break bonds /kJ mol^{-1}	Product bonds	Energy given out when bonds form /kJ mol^{-1}
4 x C—H	4 x +413 = +1652	2 x C=O	2 x —805 = —1610
2 x O=O	2 x +498 = +996	4 x O—H	4 x —464 = —1856
Total energy needed	= +2648	**Total energy given out**	= —3466

Enthalpy change of reaction = **+2648** + **(—3466)** = **—818 kJ mol^{-1}**

EXAMPLE

The information from the example before can be shown on an enthalpy level diagram.



STANDARD ENTHALPY OF ATOMISATION

Standard enthalpy of atomisation (ΔH_{atm}) of an element is the enthalpy change when **1 mole** of **gaseous atoms** is formed from the elements in their **standard states**.

$$\frac{1}{2}Cl_2(g) \longrightarrow Cl(g) \quad \Delta H = +121\ kJ\ mol^{-1}$$

For chlorine ΔH_{atm} = ½ bond dissociation energy

$$Na(s) \longrightarrow Na(g) \quad \Delta H = +108\ kJ\ mol^{-1}$$

For sodium ΔH_{atm} = energy of sublimation

STANDARD ENTHALPY OF ATOMISATION

$$\tfrac{1}{2}I_2(s) \longrightarrow I(g) \quad \Delta H = +107\ \text{kJ mol}^{-1}$$

For iodine ΔH_{atm} = ½(energy of sublimation + bond dissociation energy)

$= \tfrac{1}{2}(63 + 151) = ?$

$$\tfrac{1}{2}Br_2(l) \longrightarrow Br(g) \quad \Delta H = +112\ \text{kJ mol}^{-1}$$

For bromine ΔH_{atm} = ½(enthalpy of vaporisation + bond dissociation energy)

SKILL CHECK

The equation for the combustion of ethanol is:

$$C_2H_5OH(l) + 3O_2(g) \longrightarrow 2CO_2(g) + 3H_2O(l)$$

A. Rewrite this equation to show all the bonds in the reactants and products.

B. Use bond energies (in kJ mol^{-1}) to calculate a value for the standard enthalpy change of this reaction.

C. The standard enthalpy change of combustion of ethanol is –1367.3 kJ mol^{-1}. Suggest why this value differs from the value obtained using bond energies.

NOTES

ENERGETICS WS 1

1 For which equation is the enthalpy change correctly described as an enthalpy change of formation?

A $2NO(g) \rightarrow N_2(g) + O_2(g)$

B $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$

C $H_2O(l) + NaC\ell(s) \rightarrow NaC\ell(aq)$

D $K(s) + Mn(s) + 2O_2(g) \rightarrow KMnO_4(s)$

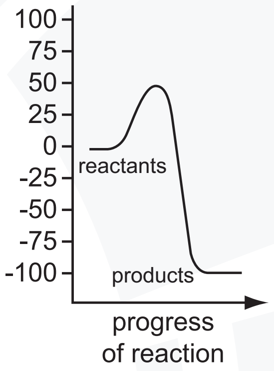
2 An exothermic chemical reaction proceeds by two stages.

$$\text{reactants} \xrightarrow{\text{stage 1}} \text{intermediate} \xrightarrow{\text{stage 2}} \text{products}$$

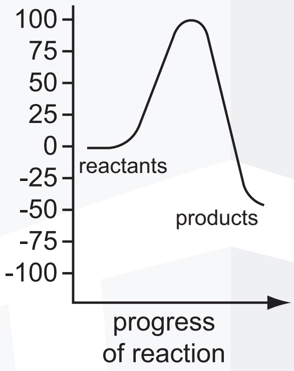
The activation energy of stage 1 is 50 kJ mol^{-1}. The overall enthalpy change of reaction is –100 kJ mol^{-1}.

Which diagram represents the reaction pathway for this reaction?

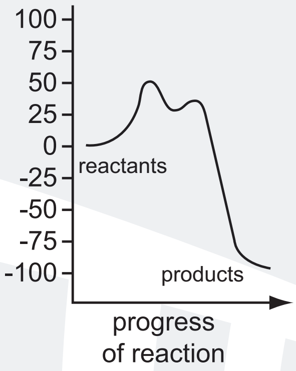
A



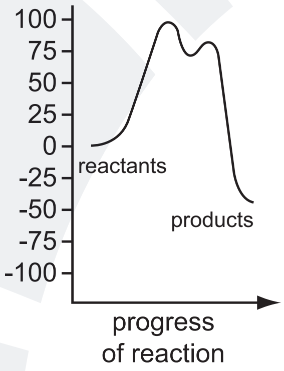
B



C



D

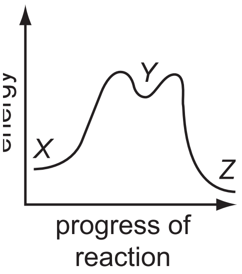


3 In the conversion of compound *X* into compound *Z*, it was found that the reaction proceeded by way of compound *Y*, which could be isolated. The following steps were involved.

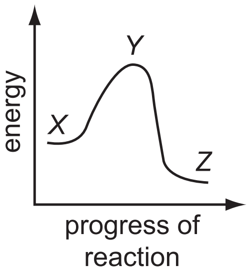
$X \rightarrow Y$; ΔH, positive
$Y \rightarrow Z$; ΔH, negative

Which reaction profile fits these data?

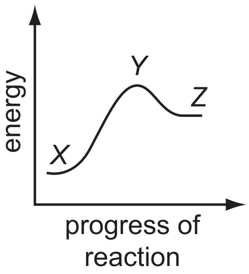
A



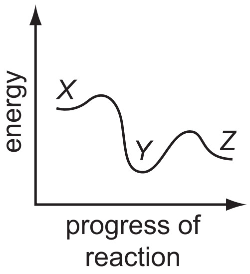
B



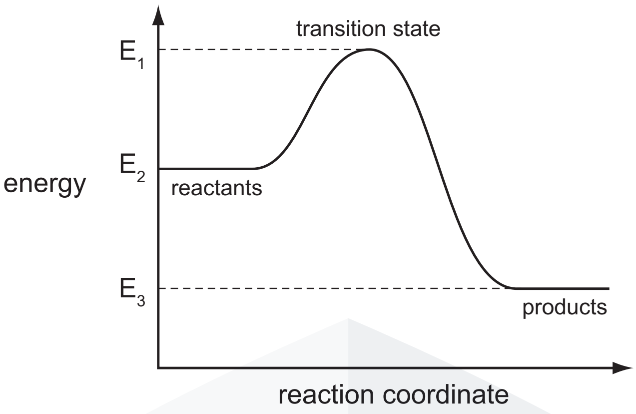
C



D



4 The reaction pathway diagram below illustrates the energies of reactants, products and the transition state of a reaction.

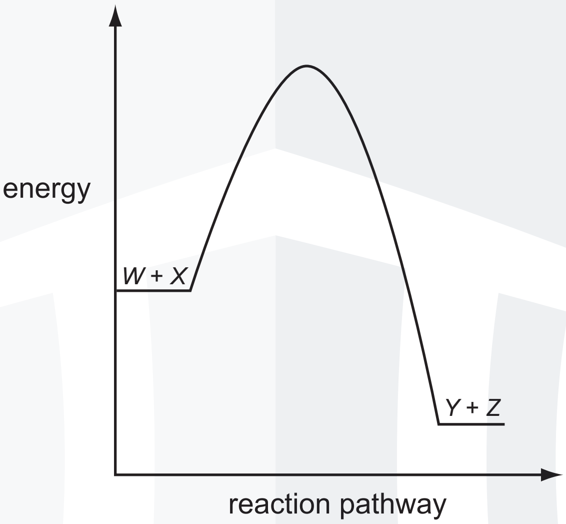


Which expression represents the activation energy of the forward reaction?

A $E_1 - E_2$ **B** $E_1 - E_3$ **C** $E_2 - E_3$ **D** $(E_1 - E_2) - (E_2 - E_3)$

5 The diagram represents the reaction pathway for the following reaction.

$$W(g) + X(g) \rightarrow Y(g) + Z(g)$$



What statement can be made about the reverse reaction, $Y(g) + Z(g) \rightarrow W(g) + X(g)$?

A It will have a larger activation energy and a positive ΔH.

B It will have a larger activation energy and a negative ΔH.

C It will have a smaller activation energy and a positive ΔH.

D It will have a smaller activation energy and a negative ΔH.

6 Which statement about the standard enthalpy change of formation of carbon dioxide is correct?

A It is equal to the standard enthalpy change of combustion of carbon.

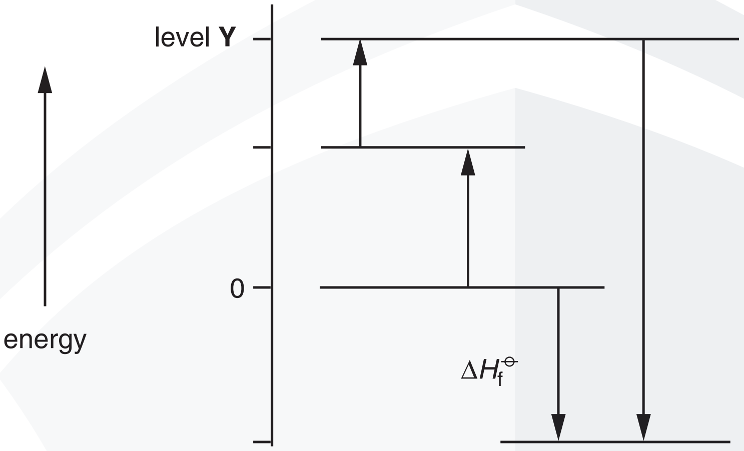
B It is equal to twice the bond energy of the C=O bond.

C It is the energy released when one mole of carbon dioxide is formed from carbon at the temperature of combustion of the carbon.

D It is the same for carbon dioxide produced from graphite and from diamond.

7 The following energy cycle represents the enthalpy changes in the formation of carbon dioxide from its constituent elements in their standard states.

What substances are present at level **Y** in this diagram?



A $C(g) + 2O(g)$

B $C(g) + O_2(g)$

C $C(s) + O_2(g)$

D $CO_2(g)$

8 50 cm^3 of 2.50 mol dm^{-3} hydrochloric acid was placed in a polystyrene beaker of negligible heat capacity. Its temperature was recorded and then 50 cm^3 of 2.50 mol dm^{-3} NaOH at the same temperature was quickly added, with stirring. The temperature rose by 17 °C.

The resulting solution may be considered to have a specific heat capacity of 4.2 J g^{-1} K^{-1}.

What is an approximate value for the molar enthalpy change of neutralisation of hydrochloric acid and sodium hydroxide from this experiment?

A $\dfrac{-(50 \times 4.2 \times 17)}{(0.050 \times 2.5)}$ J mol^{-1}

B $\dfrac{-(50 \times 4.2 \times 17)}{(0.10 \times 2.5)}$ J mol^{-1}

C $\dfrac{-(100 \times 4.2 \times 17)}{(0.050 \times 2.5)}$ J mol^{-1}

D $\dfrac{-(100 \times 4.2 \times 17)}{(50 \times 2.5)}$ J mol^{-1}

9 In a calorimetric experiment 1.60 g of a fuel is burnt. 45 % of the energy released is absorbed by 200 g of water whose temperature rises from 18 °C to 66 °C. The specific heat capacity of water is 4.2 J g^{-1} K^{-1}.

What is the total energy released per gram of fuel burnt?

A 25 200 J **B** 56 000 J **C** 89 600 J **D** 143 360 J

10 The equation below represents the combination of gaseous atoms of non-metal X and of hydrogen to form gaseous X_2H_6 molecules.

$$2X(g) + 6H(g) \rightarrow X_2H_6(g) \qquad \Delta H = -2775\ \text{kJ mol}^{-1}$$

The bond energy of an X–H bond is 395 kJ mol^{-1}.

What is the bond energy of an X–X bond?

A –405.0 kJ mol^{-1}

B –202.5 kJ mol^{-1}

C +202.5 kJ mol^{-1}

D +405.0 kJ mol^{-1}

11 Methanol may be prepared by the reaction between carbon monoxide and hydrogen.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$$

The relevant average bond energies are given below.

E(C≡O)	1077 kJ mol^{-1}
E(C–O)	360 kJ mol^{-1}
E(C–H)	410 kJ mol^{-1}
E(H–H)	436 kJ mol^{-1}
E(O–H)	460 kJ mol^{-1}

What is the enthalpy change of this reaction?

A –537 kJ mol^{-1}

B –101 kJ mol^{-1}

C +101 kJ mol^{-1}

D +537 kJ mol^{-1}

12 Which equation represents the change corresponding to the enthalpy change of atomisation of iodine?

A $\frac{1}{2}I_2(g) \rightarrow I(g)$

B $I_2(g) \rightarrow 2I(g)$

C $\frac{1}{2}I_2(s) \rightarrow I(g)$

D $I_2(s) \rightarrow 2I(g)$

13 Some bond energy values are listed below.

bond	bond energy / kJ mol^{-1}
C–H	410
C–Cl	340
Cl–Cl	244
Br–Br	193

These bond energy values relate to the following four reactions.

P $Br_2 \rightarrow 2Br$

Q $2Cl \rightarrow Cl_2$

R $CH_3 + Cl \rightarrow CH_3Cl$

S $CH_4 \rightarrow CH_3 + H$

What is the order of enthalpy changes of these reactions from most negative to most positive?

A P → Q → R → S

B Q → R → S → P

C R → Q → P → S

D S → P → Q → R

14 Which equation represents the standard enthalpy change of atomisation of bromine?

A $Br_2(l) \rightarrow 2Br(g)$

B $Br_2(g) \rightarrow 2Br(g)$

C $\frac{1}{2}Br_2(l) \rightarrow Br(g)$

D $\frac{1}{2}Br_2(g) \rightarrow Br(g)$

15 The gaseous oxides of nitrogen have positive enthalpy changes of formation.

Which factor is likely to make the most significant contribution to these enthalpy changes?

A the high bond energy of the nitrogen molecule, N_2

B the high electron affinity of nitrogen atoms

C the high electron affinity of oxygen atoms

D the similarity of the electronegativities of oxygen and nitrogen

16 Hydrazine, N_2H_4, is widely used as a rocket fuel because it reacts with oxygen as shown, producing ‘environmentally friendly’ gases.

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g) \qquad \Delta H = -534\,\text{kJ mol}^{-1}$$

Despite its use as a rocket fuel, hydrazine does not burn spontaneously in oxygen.

Which statement explains why hydrazine does **not** burn spontaneously?

A Hydrazine is a liquid.

B The activation energy is too high.

C The $N\equiv N$ bond is very strong.

D The reaction is exothermic.

17 Skiers trapped by snowstorms use heat packs to keep warm. The heat may be generated by the reaction below.

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s) \quad ; \quad \Delta H^{\ominus} = -1648\,\text{kJ}$$

What is the standard enthalpy change of formation of iron(III) oxide?

A $0\,\text{kJ mol}^{-1}$

B $-824\,\text{kJ mol}^{-1}$

C $-1648\,\text{kJ mol}^{-1}$

D $-3296\,\text{kJ mol}^{-1}$

18 At 600 °C oxides of nitrogen react with unburnt hydrocarbons in a catalytic converter in a car exhaust. The equation using methane as representative of a hydrocarbon molecule would be as follows.

$$4NO + CH_4 \rightarrow 2N_2 + CO_2 + 2H_2O$$

Which statement is likely to be true about the energy change of this reaction?

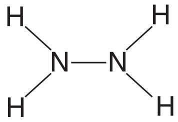
A It is endothermic as heat energy is converted into chemical energy.

B It is exothermic as a high temperature is required.

C It is exothermic as the triple bond $N\underset{\leftarrow}{=}O$ is broken.

D It is exothermic as the products have large negative enthalpy changes of formation.

19 Hydrazine was used as a fuel for the Messerschmidt 163 rocket fighter in World War II and for the American Gemini and Apollo spacecraft. It has the following formula.



What is the enthalpy change of atomisation of 1 mol of gaseous hydrazine?

A 550 kJ

B 1720 kJ

C 1970 kJ

D 2554 kJ

20 The 'flash' produced by nineteenth century photographers to take indoor photographs was obtained from the following reaction.

$$3Mg + KClO_3 \longrightarrow 3MgO + KCl$$

The standard enthalpy changes of formation are given below.

	$\Delta H_f^\ominus$ / kJ mol^{-1}
MgO	–602
KC*l*	–437
$KClO_3$	–391

What is the standard enthalpy change of the 'flash' reaction?

A –3(–602) + (–437) – (–391)

B (–602) + (–437) – (–391)

C 3(–602) + (–437) – (–391)

D (–602) + 3(–437) – 3(–391)

21 Titanium occurs naturally as the mineral rutile, TiO_2. One possible method of extraction of titanium is to reduce the rutile by heating with carbon.

$$TiO_2(s) + 2C(s) \rightarrow Ti(s) + 2CO(g)$$

The standard enthalpy changes of formation of TiO_2(s) and CO(g) are –940 kJ mol^{-1} and –110 kJ mol^{-1} respectively.

What is the standard enthalpy change of this reaction?

A –830 kJ mol^{-1}

B –720 kJ mol^{-1}

C +720 kJ mol^{-1}

D +830 kJ mol^{-1}

22 Hydrogen peroxide slowly decomposes into water and oxygen. The enthalpy change of reaction can be calculated using standard enthalpies of formation.

$\Delta H_f^\ominus$(hydrogen peroxide(l)) = –187.8 kJ mol^{-1}

$\Delta H_f^\ominus$(water(l)) = –285.8 kJ mol^{-1}

Using a Hess cycle, what is the enthalpy change of reaction for this decomposition?

$$2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$$

A +98 kJ mol^{-1}

B –98 kJ mol^{-1}

C –196 kJ mol^{-1}

D –947.2 kJ mol^{-1}

23 The first stage in the industrial production of nitric acid from ammonia can be represented by the following equation.

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$$

Using the following standard enthalpy change of formation data, what is the value of the standard enthalpy change, $\Delta H^\ominus$, for this reaction?

compound	$\Delta H_f^\ominus$ / kJ mol^{-1}
$NH_3(g)$	–46.1
NO(g)	+90.3
$H_2O(g)$	–241.8

A +905.2 kJ mol^{-1}

B –105.4 kJ mol^{-1}

C –905.2 kJ mol^{-1}

D –1274.0 kJ mol^{-1}

24 Given $S(s) + O_2(g) \rightarrow SO_2(g)$, $\Delta H_f^\ominus = -297$ kJ mol^{-1}
and $S(s) + 1\frac{1}{2}O_2(g) \rightarrow SO_3(g)$ $\Delta H_f^\ominus = -395$ kJ mol^{-1}

what is the enthalpy change of reaction, $\Delta H^\ominus$, of $2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$?

A –196 kJ mol^{-1} **B** –98 kJ mol^{-1} **C** +98 kJ mol^{-1} **D** +196 kJ mol^{-1}

25 The standard enthalpy changes of formation of iron(II) oxide, FeO(s), and aluminium oxide, $Al_2O_3(s)$, are –266 kJ mol^{-1} and –1676 kJ mol^{-1} respectively.

What is the enthalpy change under standard conditions for the following reaction?

$$3FeO(s) + 2Al(s) \rightarrow 3Fe(s) + Al_2O_3(s)$$

A + 878 kJ **B** – 878 kJ **C** –1942 kJ **D** –2474 kJ

26 Why does the exothermic reaction

$$C(diamond) \longrightarrow C(graphite) \qquad \Delta H = -3\,kJ\,mol^{-1}$$

not occur spontaneously?

A A tetrahedral configuration is always more stable than a planar one.

B Diamond has only strong covalent bonds whereas graphite has both covalent bonds and van der Waals' forces.

C The change from diamond to graphite has a high activation energy.

D The density of graphite is less than that of diamond.

27 Enthalpy changes of combustion can be used to determine enthalpy changes of formation. The following equation represents the enthalpy change of formation of butane.

$$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g)$$

By using the following standard enthalpy of combustion data, what is the value of the standard enthalpy change of formation, $\Delta H^{\ominus}_{f}$, for this reaction?

compound	$\Delta H^{\ominus}_{c}$ / kJ mol^{-1}
carbon	–394
hydrogen	–286
butane	–2877

A –5883 kJ mol^{-1}

B –129 kJ mol^{-1}

C +129 kJ mol^{-1}

D +2197 kJ mol^{-1}

28 Which of these compounds, on complete combustion of one mole, will release the greatest amount of energy?

A $CH_3CH_2CH_2OH$

B $CH_3CH_2CH_3$

C $CH_3CH_2CO_2H$

D CH_3COCH_3

29 Given the following enthalpy changes,

$$I_2(g) + 3Cl_2(g) \rightarrow 2ICl_3(s) \qquad \Delta H^\ominus = -214\,kJ\,mol^{-1}$$

$$I_2(s) \rightarrow I_2(g) \qquad \Delta H^\ominus = +38\,kJ\,mol^{-1}$$

What is the standard enthalpy change of formation of iodine trichloride, $ICl_3(s)$?

A $+176\,kJ\,mol^{-1}$

B $-88\,kJ\,mol^{-1}$

C $-176\,kJ\,mol^{-1}$

D $-214\,kJ\,mol^{-1}$

30 Iodine trichloride, ICl_3, is made by reacting iodine with chlorine.

$$I_2(s) + Cl_2(g) \rightarrow 2ICl(s)\ ;\ \Delta H^\ominus = +14\,kJ\,mol^{-1}$$

$$ICl(s) + Cl_2(g) \rightarrow ICl_3(s)\ ;\ \Delta H^\ominus = -88\,kJ\,mol^{-1}$$

By using the data above, what is the enthalpy change of the formation for solid iodine trichloride?

A $-60\,kJ\,mol^{-1}$

B $-74\,kJ\,mol^{-1}$

C $-81\,kJ\,mol^{-1}$

D $-162\,kJ\,mol^{-1}$

31 The enthalpy change of formation, ΔH_f, of hydrated calcium ions is the enthalpy change of the following reaction.

$$Ca(s) + aq - 2e^- \rightarrow Ca^{2+}(aq)$$

The following enthalpy changes are **not** quoted in the *Data Booklet*.

$$Ca(s) \rightarrow Ca(g) \qquad \Delta H_a = 177\,kJ\,mol^{-1}$$

$$Ca^{2+}(g) + aq \rightarrow Ca^{2+}(aq) \qquad \Delta H_{hyd} = -1565\,kJ\,mol^{-1}$$

What is the enthalpy change of formation of hydrated calcium ions?

A $-1388\,kJ\,mol^{-1}$

B $-798\,kJ\,mol^{-1}$

C $-238\,kJ\,mol^{-1}$

D $+352\,kJ\,mol^{-1}$

32 Given

$CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g) \quad \Delta H^\ominus = -283\ \text{kJ mol}^{-1}$

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l) \quad \Delta H^\ominus = -286\ \text{kJ mol}^{-1}$

$H_2O(g) \rightarrow H_2O(l) \quad \Delta H^\ominus = -44\ \text{kJ mol}^{-1}$

what is the change in enthalpy, $\Delta H^\ominus$, for the following reaction?

$$CO_2(g) + H_2(g) \rightarrow CO(g) + H_2O(g)$$

A $-525\ \text{kJ mol}^{-1}$ **B** $-41\ \text{kJ mol}^{-1}$ **C** $+41\ \text{kJ mol}^{-1}$ **D** $+525\ \text{kJ mol}^{-1}$

33 Red lead oxide, Pb_3O_4, is used in metal priming paints. It can be made by heating PbO in air.

$$6PbO(s) + O_2(g) \rightarrow 2Pb_3O_4(s)$$

Which two values are needed to calculate the enthalpy change for this reaction?

A enthalpy change of combustion of lead and enthalpy change of formation of Pb_3O_4

B enthalpy change of combustion of PbO and enthalpy change of formation of Pb_3O_4

C enthalpy change of formation of PbO and enthalpy change of atomisation of O_2

D enthalpy change of formation of PbO and enthalpy change of formation of Pb_3O_4

34 The standard enthalpy changes of formation of HCl and HI are $-92\ \text{kJ mol}^{-1}$ and $+26\ \text{kJ mol}^{-1}$ respectively.

Which statement is **most** important in explaining this difference?

A Chlorine is more electronegative than iodine.

B The activation energy for the H_2/Cl_2 reaction is much less than that for the H_2/I_2 reaction.

C The bond energy of HI is smaller than the bond energy of HCl.

D The bond energy of I_2 is smaller than the bond energy of Cl_2.

35 The standard enthalpy change for the reaction

$$2NF_3(g) \rightarrow 2N(g) + 6F(g) \qquad \text{is } \Delta H^\ominus = +1668\ \text{kJ}$$

What is the bond energy of the N–F bond?

A $-556\ \text{kJ mol}^{-1}$

B $-278\ \text{kJ mol}^{-1}$

C $+278\ \text{kJ mol}^{-1}$

D $+556\ \text{kJ mol}^{-1}$

36 Which reaction is endothermic?

A $2HBr \rightarrow H_2 + Br_2$

B $N_2 + 3H_2 \rightarrow 2NH_3$

C $2SO_2 + O_2 \rightarrow 2SO_3$

D $SO_3 + H_2O \rightarrow H_2SO_4$

37 Hess's Law can be used to calculate the average C-H bond energy in methane.

$\Delta H^\ominus_{at}$ = standard enthalpy change of atomisation

$\Delta H^\ominus_{f}$ = standard enthalpy change of formation

$\Delta H^\ominus_{c}$ = standard enthalpy change of combustion

Which data values are needed in order to perform the calculation?

A $\Delta H^\ominus_{at}$ (C), $\Delta H^\ominus_{at}$ (H), $\Delta H^\ominus_{f}$ (CH_4)

B $\Delta H^\ominus_{c}$ (C), $\Delta H^\ominus_{c}$ (H_2), $\Delta H^\ominus_{c}$ (CH_4)

C $\Delta H^\ominus_{c}$ (C), $\Delta H^\ominus_{c}$ (H_2), $\Delta H^\ominus_{f}$ (CH_4)

D $\Delta H^\ominus_{f}$ (CH_4) only, as $\Delta H^\ominus_{f}$ (C), and $\Delta H^\ominus_{f}$ (H_2), are defined as zero

38 The standard enthalpy changes of formation of HCl and HI are –92 kJ mol^{-1} and +26 kJ mol^{-1} respectively.

Which statement is **most** important in explaining this difference?

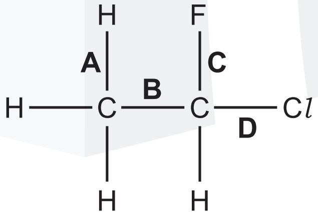
A Chlorine is more electronegative than iodine.

B The activation energy for the $H_2 + Cl_2$ reaction is much less than that for the $H_2 + I_2$ reaction.

C The bond energy of HI is smaller than the bond energy of HCl.

D The bond energy of I_2 is smaller than the bond energy of Cl_2.

39 Which bond in the structure below has the lowest bond energy?



40 The equation for a reaction is shown.

$$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)\ ;\ \Delta H = x\ \text{kJ mol}^{-1}$$

Which pair of descriptions is fully correct for this reaction?

	type(s) of enthalpy change	value of *x*
A	formation only	positive
B	formation only	negative
C	combustion, formation	positive
D	combustion, formation	negative

41 Propanone has the molecular formula C_3H_6O.

The enthalpy change of combustion of hydrogen is $-286\ \text{kJ mol}^{-1}$.
The enthalpy change of combustion of carbon is $-394\ \text{kJ mol}^{-1}$.
The enthalpy change of combustion of propanone is $-1786\ \text{kJ mol}^{-1}$.

Using this information, what is the enthalpy change of formation of propanone?

A $-1106\ \text{kJ mol}^{-1}$

B $-540\ \text{kJ mol}^{-1}$

C $-254\ \text{kJ mol}^{-1}$

D $+1106\ \text{kJ mol}^{-1}$

42 A reaction which causes the presence of oxides of nitrogen in car exhausts is the formation of NO.

$$N_2 + O_2 \rightarrow 2NO \qquad \Delta H = +180\ \text{kJ mol}^{-1}$$

What is the bond energy in kJ mol^{-1} of the bond between the atoms in NO?

A 655 **B** 835 **C** 1310 **D** 1670

43 In the table below,

- '+' means that this type of standard enthalpy change can **only** have positive values,
- '–' means that this type of standard enthalpy change can **only** have negative values,
- '+/–' means that **either** positive **or** negative values are possible.

Which row is correct?

	atomisation	formation	solution
A	+	+	+/–
B	+	+/–	+/–
C	–	+/–	–
D	–	–	+

44 A student calculated the standard enthalpy change of formation of ethane, C_2H_6, using a method based on standard enthalpy changes of combustion.

He used correct values for the standard enthalpy change of combustion of ethane (–1560 kJ mol^{-1}) and hydrogen (–286 kJ mol^{-1}) but he used an incorrect value for the standard enthalpy change of combustion of carbon. He then performed his calculation correctly. His final answer was –158 kJ mol^{-1}.

What did he use for the standard enthalpy change of combustion of carbon?

A –1432 kJ mol^{-1}

B –860 kJ mol^{-1}

C –430 kJ mol^{-1}

D –272 kJ mol^{-1}

45 Which process could be used to calculate the bond energy for the covalent bond X-Y by dividing its ΔH by n?

A $XY_n(g) \rightarrow X(g) + nY(g)$

B $2XY_n(g) \rightarrow 2XY_{n-1}(g) + Y_2(g)$

C $Y(g) + XY_{n-1}(g) \rightarrow XY_n(g)$

D $nXY(g) \rightarrow nX(g) + \frac{n}{2}Y_2(g)$

46 The enthalpy change of the neutralisation given below is –114 kJ mol^{-1}.

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

By using this information, what is the most likely value for the enthalpy change of the following neutralisation?

$$Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l)$$

A –57 kJ mol^{-1} **B** –76 kJ mol^{-1} **C** –114 kJ mol^{-1} **D** –228 kJ mol^{-1}

47 Ethyne, C_2H_2, completely combusts, as shown in the equation.

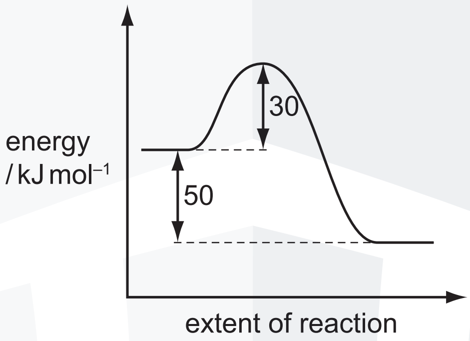
$$\mathrm{H{-}C{\equiv}C{-}H} + 2\tfrac{1}{2}\,\mathrm{O{=}O} \rightarrow \mathrm{H_2O} + 2\mathrm{CO_2}$$

Using the average bond enthalpies in the table, what is the enthalpy change of combustion of ethyne?

bond	average bond enthalpy / $kJ\,mol^{-1}$
C—H	410
C≡C	840
O═O	496
C═O	740
O—H	460
C—O	360

A $-980\ kJ\,mol^{-1}$ **B** $-540\ kJ\,mol^{-1}$ **C** $+540\ kJ\,mol^{-1}$ **D** $+980\ kJ\,mol^{-1}$

48 The reaction pathway for a reversible reaction is shown below.



Which statement is correct?

A The activation energy of the reverse reaction is $+80\ kJ\,mol^{-1}$.

B The enthalpy change for the forward reaction is $+30\ kJ\,mol^{-1}$.

C The enthalpy change for the forward reaction is $+50\ kJ\,mol^{-1}$.

D The enthalpy change for the reverse reaction is $+30\ kJ\,mol^{-1}$.

49 This question should be answered using bond enthalpy data. The equation for the complete combustion of methane is given below.

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

What is the enthalpy change of combustion of methane?

A $-1530\,kJ\,mol^{-1}$

B $-1184\,kJ\,mol^{-1}$

C $-770\,kJ\,mol^{-1}$

D $-688\,kJ\,mol^{-1}$

50 A student mixed $25\,cm^3$ of $0.10\,mol\,dm^{-3}$ sodium hydroxide solution with $25\,cm^3$ of $0.10\,mol\,dm^{-3}$ hydrochloric acid and noted a temperature rise of 2.5 °C.

What is the enthalpy change of the reaction per mole of NaOH?

A $-209\,kJ\,mol^{-1}$

B $-104.5\,kJ\,mol^{-1}$

C $-209\,J\,mol^{-1}$

D $-522.5\,J\,mol^{-1}$

51 Which energy change corresponds to the enthalpy change of atomisation of hydrogen at 298 K?

A the bond energy of a H–H bond

B half the bond energy of a H–H bond

C minus half the bond energy of a H–H bond

D minus the bond energy of a H–H bond

52 Propanone has molecular formula C_3H_6O.

The enthalpy change of combustion of hydrogen is $-286\,kJ\,mol^{-1}$.

The enthalpy change of combustion of carbon is $-394\,kJ\,mol^{-1}$.

The enthalpy change of formation of propanone is $-254\,kJ\,mol^{-1}$.

Using this information, what is the enthalpy change of combustion of propanone?

A $-2644\,kJ\,mol^{-1}$

B $-2294\,kJ\,mol^{-1}$

C $-1786\,kJ\,mol^{-1}$

D $-426\,kJ\,mol^{-1}$

53 This question should be answered using bond enthalpy data. The equation for the complete combustion of methanal is given below.

$$H_2C{=}O + O_2 \rightarrow CO_2 + H_2O$$

What is the enthalpy change of combustion of methanal?

A $+416\,\text{kJ mol}^{-1}$

B $+396\,\text{kJ mol}^{-1}$

C $-344\,\text{kJ mol}^{-1}$

D $-690\,\text{kJ mol}^{-1}$

54 Equations involving four enthalpy changes are shown.

$Na(g) \rightarrow Na^+(g) + e^-$ $\Delta H = W$

$Na(g) \rightarrow Na^{2+}(g) + 2e^-$ $\Delta H = X$

$Na(s) \rightarrow Na(g)$ $\Delta H = Y$

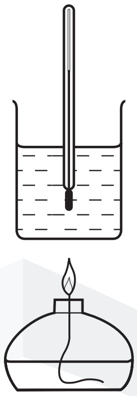
$Na(s) \rightarrow Na^{2+}(g) + 2e^-$ $\Delta H = Z$

What is the second ionisation energy of sodium?

A 2W **B** X – W **C** Y – W **D** Z – Y

55 A student carried out an experiment to determine the enthalpy change for the combustion of methanol.

The following results were obtained by the student.



start temperature of the water	20 °C
final temperature of the water	53 °C
mass of alcohol burner before burning	259.65 g
mass of alcohol burner after burning	259.15 g
mass of glass beaker plus water	150.00 g
mass of glass beaker	50.00 g

How much of the heat energy produced by the burning of methanol went into the water?

A 209 J **B** 13 794 J **C** 20 691 J **D** 22 154 J

56 A student mixed 25.0 cm^3 of 0.350 mol dm^{-3} sodium hydroxide solution with 25.0 cm^3 of 0.350 mol dm^{-3} hydrochloric acid. The temperature rose by 2.50 °C. Assume that no heat was lost to the surroundings.

The final mixture had a specific heat capacity of 4.20 J cm^{-3} K^{-1}.

What is the molar enthalpy change for the reaction?

A –150 kJ mol^{-1}

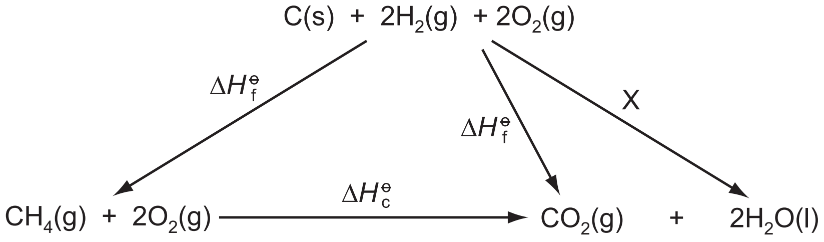
B –60.0 kJ mol^{-1}

C –30.0 kJ mol^{-1}

D –0.150 kJ mol^{-1}

57 Enthalpy changes that are difficult to measure directly can often be determined using Hess' Law to construct an enthalpy cycle.

Which enthalpy change is indicated by X in the enthalpy cycle shown?



A $-4 \times$ the enthalpy of combustion of hydrogen

B $+4 \times$ the enthalpy of combustion of hydrogen

C $-2 \times$ the enthalpy of formation of water

D $+2 \times$ the enthalpy of formation of water

58 The enthalpy change of formation of carbon dioxide is $-394\,\text{kJ mol}^{-1}$.
The enthalpy change of formation of water is $-286\,\text{kJ mol}^{-1}$.
The enthalpy change of formation of methane is $-74\,\text{kJ mol}^{-1}$.

What is the enthalpy change of combustion of methane?

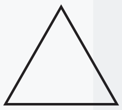
A $-892\,\text{kJ mol}^{-1}$

B $-606\,\text{kJ mol}^{-1}$

C $+606\,\text{kJ mol}^{-1}$

D $+892\,\text{kJ mol}^{-1}$

59 The diagram shows the skeletal formula of cyclopropane.



The enthalpy change of formation of cyclopropane is $+53.3\,\text{kJ mol}^{-1}$ and the enthalpy change of atomisation of graphite is $+717\,\text{kJ mol}^{-1}$.

The bond enthalpy of H–H is $436\,\text{kJ mol}^{-1}$ and of C–H is $410\,\text{kJ mol}^{-1}$.

What value for the average bond enthalpy of the C–C bond in cyclopropane can be calculated from this data?

A $187\,\text{kJ mol}^{-1}$ **B** $315\,\text{kJ mol}^{-1}$ **C** $351\,\text{kJ mol}^{-1}$ **D** $946\,\text{kJ mol}^{-1}$

60 The enthalpy change of formation of $Mn(NO_3)_2(s)$ is –696 kJ mol^{-1}.
The enthalpy change of formation of $MnO_2(s)$ is –520 kJ mol^{-1}.
The enthalpy change of formation of $NO_2(g)$ is +33 kJ mol^{-1}.

On heating, $Mn(NO_3)_2$ decomposes into MnO_2 and NO_2.

$$Mn(NO_3)_2(s) \rightarrow MnO_2(s) + 2NO_2(g)$$

What is the value of the standard enthalpy change of this reaction?

A –242 kJ mol^{-1}

B –209 kJ mol^{-1}

C +209 kJ mol^{-1}

D +242 kJ mol^{-1}

61 Which reaction has an enthalpy change equal to the standard enthalpy change of formation of propane?

A $3C(g) + 4H_2(g) \rightarrow C_3H_8(g)$

B $3C(g) + 8H(g) \rightarrow C_3H_8(g)$

C $3C(s) + 4H_2(g) \rightarrow C_3H_8(g)$

D $3C(s) + 4H_2(g) \rightarrow C_3H_8(l)$

62 A student investigates four different fuels. Each fuel is used separately to raise the temperature of 1 dm^3 of water from 20 °C to 100 °C. Each fuel undergoes complete combustion. All other conditions are the same in each experiment.

Which fuel would produce the **smallest** amount of carbon dioxide in these experiments?

	fuel	energy released per mole of fuel
A	ethanol	1367 kJ mol^{-1}
B	methane	890 kJ mol^{-1}
C	methanol	715 kJ mol^{-1}
D	propane	2220 kJ mol^{-1}

63 Ethanol is increasingly being used as a fuel for cars.

The standard enthalpy change of formation of carbon dioxide is –393 kJ mol^{-1}.
The standard enthalpy change of formation of water is –286 kJ mol^{-1}.
The standard enthalpy change of formation of ethanol is –277 kJ mol^{-1}.

What is the standard enthalpy change of combustion of ethanol?

A –1921 kJ mol^{-1}

B –1367 kJ mol^{-1}

C –956 kJ mol^{-1}

D –402 kJ mol^{-1}

64 In an experiment to calculate the enthalpy change of combustion of a fuel, 1.5 g (0.0326 mol) of the fuel was used to heat 200 g of water. The temperature of the water rose from 25 °C to 55 °C. The specific heat capacity of water is 4.18 J g^{-1} K^{-1}.

There is significant heat loss in this experiment. Therefore, the experimental value for the enthalpy change of combustion, ΔH_c, of the fuel will be different from the theoretical value.

Using the information above, what is the experimental value for the enthalpy change of combustion, ΔH_c, of the fuel?

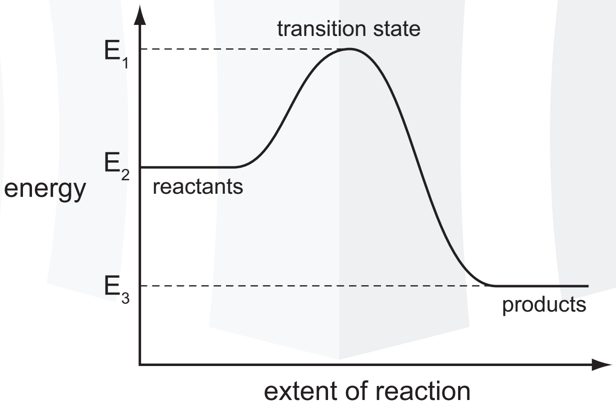
A –1410 kJ mol^{-1}

B –769 kJ mol^{-1}

C –30.7 kJ mol^{-1}

D –16.7 kJ mol^{-1}

65 The reaction pathway diagram below illustrates the energies of the reactants, the products and the transition state of a reaction.



Which expression represents the activation energy of the forward reaction?

A $E_1 - E_2$ **B** $E_2 - E_1$ **C** $E_2 - E_3$ **D** $E_3 - E_2$

66 For which equation is the enthalpy change correctly described as an enthalpy change of formation?

A $C(g) + O_2(g) \rightarrow CO_2(g)$

B $C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$

C $2N(g) + 4O(g) \rightarrow N_2O_4(g)$

D $2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

67 Hydrogen sulfide, H_2S, is released from volcanoes. It reacts with oxygen in the air to form sulfur dioxide.

$$2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g)$$

$\Delta H_f^\ominus [H_2S(g)] = -21\,kJ\,mol^{-1}$

$\Delta H_f^\ominus [H_2O(l)] = -286\,kJ\,mol^{-1}$

$\Delta H_f^\ominus [SO_2(g)] = -297\,kJ\,mol^{-1}$

What is the standard enthalpy change of this reaction?

A $-1208\,kJ\,mol^{-1}$

B $-1124\,kJ\,mol^{-1}$

C $-562\,kJ\,mol^{-1}$

D $-541\,kJ\,mol^{-1}$

68 When 0.47 g of a hydrocarbon was completely burnt in air, the energy released heated 200 g of water from 23.7 °C to 41.0 °C.

What was the amount of energy absorbed by the water?

A $0.47 \times 4.18 \times 17.3\,J$

B $0.47 \times 4.18 \times (273 + 17.3)\,J$

C $200 \times 4.18 \times 17.3\,J$

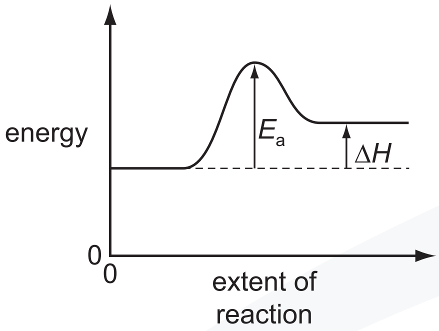
D $200 \times 4.18 \times (273 + 17.3)\,J$

69 Nitrogen monoxide is an atmospheric pollutant that is formed inside car engines by an endothermic reaction between nitrogen and oxygen.

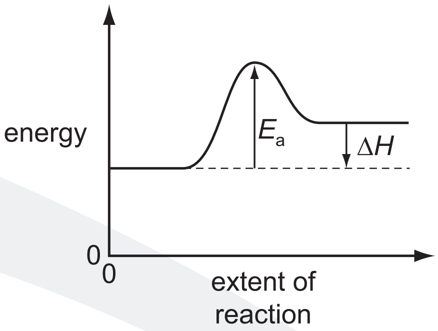
$$N_2(g) + O_2(g) \rightarrow 2NO(g) \qquad \Delta H = +66\,\text{kJ mol}^{-1}$$

Which labelled diagram correctly represents the energy profile for this reaction?

A



B



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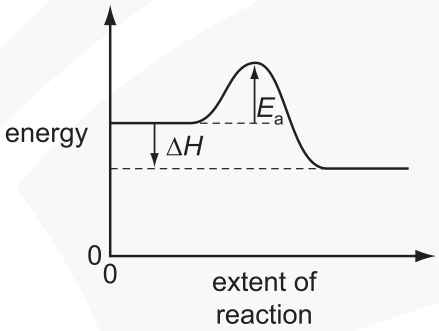

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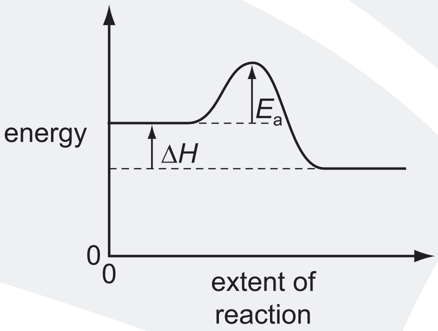


70 The standard enthalpy changes of combustion of glucose and ethanol are given as –2820 and –1368 kJ mol^{-1} respectively.

Glucose, $C_6H_{12}O_6$, can be converted into ethanol.

$$C_6H_{12}O_6(s) \rightarrow 2C_2H_5OH(l) + 2CO_2(g)$$

What is the standard enthalpy change for this reaction?

A –1452 kJ mol^{-1}

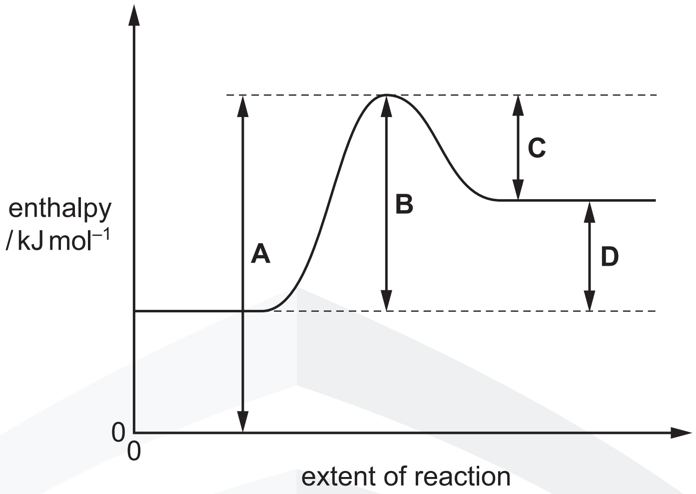
B –84 kJ mol^{-1}

C +84 kJ mol^{-1}

D +1452 kJ mol^{-1}

71 The diagram shows a reaction pathway for an endothermic reaction.

Which arrow represents the activation energy for the forward reaction?



72 Metaldehyde, $(CH_3CHO)_4$, is used as a solid fuel for camping stoves. The equation for the complete combustion of metaldehyde is shown.

$$(CH_3CHO)_4(s) + 10O_2(g) \rightarrow 8CO_2(g) + 8H_2O(l)$$

ΔH°_c = standard enthalpy change of combustion.

Which expression will give a correct value for the enthalpy change of formation of metaldehyde?

A ΔH°_c metaldehyde – ($8\Delta H^\circ_c$ carbon + $8\Delta H^\circ_c$ hydrogen)

B ΔH°_c metaldehyde – ($8\Delta H^\circ_c$ carbon + $16\Delta H^\circ_c$ hydrogen)

C ($8\Delta H^\circ_c$ carbon + $8\Delta H^\circ_c$ hydrogen) – ΔH°_c metaldehyde

D ($8\Delta H^\circ_c$ carbon + $16\Delta H^\circ_c$ hydrogen) – ΔH°_c metaldehyde

73 A student performed an experiment to measure the enthalpy change of combustion of ethane.

He used the following values for the standard enthalpy changes of combustion of carbon and hydrogen.

ΔH°_c carbon = –394 kJ mol^{-1}

ΔH°_c hydrogen = –286 kJ mol^{-1}

He calculated the enthalpy change of formation of ethane to be –140 kJ mol^{-1}.

What was his experimental value for the standard enthalpy change of combustion of ethane?

A –2364 kJ mol^{-1}

B –1506 kJ mol^{-1}

C –1112 kJ mol^{-1}

D –540 kJ mol^{-1}

74 Carbon monoxide and methanol can react together to form ethanoic acid.

$$CO(g) + CH_3OH(l) \xrightarrow{\Delta H_r^\ominus} CH_3CO_2H(l)$$

Standard enthalpy changes of combustion are given in the table.

compound	standard enthalpy change of combustion, $\Delta H_c^\ominus$
CO	–283.0 kJ mol^{-1}
CH_3OH	–726.0 kJ mol^{-1}
CH_3CO_2H	–874.1 kJ mol^{-1}

What is the value for $\Delta H_r^\ominus$ for the reaction between carbon monoxide and methanol?

A –1883.1 kJ mol^{-1}

B –134.9 kJ mol^{-1}

C +134.9 kJ mol^{-1}

D +1883.1 kJ mol^{-1}

75 Solid sulfur consists of molecules made up of eight atoms covalently bonded together.

The bonding in sulfur dioxide is O=S=O.

enthalpy change of combustion of S_8, $\Delta H_c^\ominus S_8(s)$= –2376kJ mol^{-1}
energy required to break 1 mole $S_8(s)$ into gaseous atoms = 2232 kJ mol^{-1}
O=O bond enthalpy = 496 kJ mol^{-1}

Using these data, what is the value of the S=O bond enthalpy?

A 239 kJ mol^{-1} **B** 257 kJ mol^{-1} **C** 319 kJ mol^{-1} **D** 536 kJ mol^{-1}

76 In an experiment, the burning of 1.45 g (0.025 mol) of propanone was used to heat 100 g of water. The initial temperature of the water was 20.0 °C and the final temperature of the water was 78.0 °C.

Which experimental value for the enthalpy change of combustion for propanone can be calculated from these results?

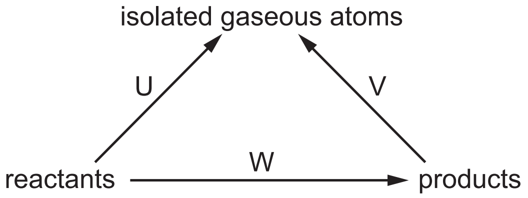
A –1304 kJ mol^{-1}

B –970 kJ mol^{-1}

C –352 kJ mol^{-1}

D –24.2 kJ mol^{-1}

77 Hess' law may be used to determine enthalpy changes using average bond energies, as shown in the diagram.



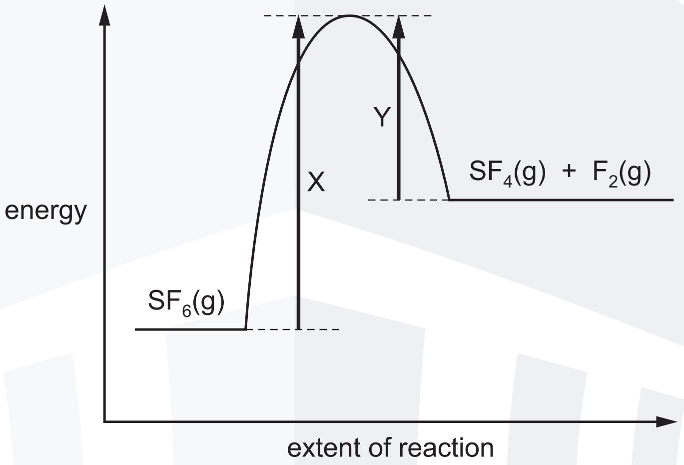
U is the sum of the average bond energies of the reactants, and V is the sum of the average bond energies of the products.

For the reaction shown below, which expression will give a value for W, the enthalpy change of combustion of methane?

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$$

A U – V **B** U + V **C** 2(U – V) **D** V – U

78 The decomposition reaction $SF_6(g) \rightarrow SF_4(g) + F_2(g)$ can be described by the reaction pathway diagram shown.



What are the values of $\Delta H^{\ominus}$ and E_a for this reaction?

	$\Delta H^{\ominus}$	E_a
A	X	X + Y
B	X	Y
C	X – Y	X
D	Y – X	X

79 Which row correctly describes what happens when the temperature of a chemical reaction is **decreased**?

	activation energy (E_a)	number of successful collisions
A	decreases	decreases
B	decreases	increases
C	remains the same	decreases
D	remains the same	increases

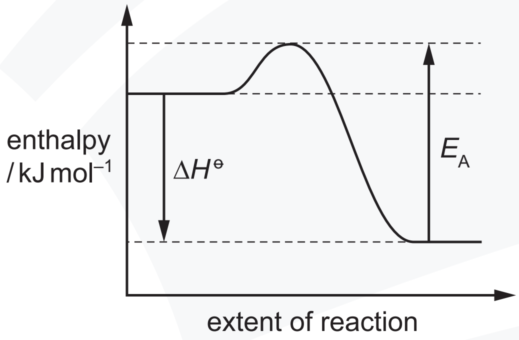
80 Nitric oxide, NO, and bromine vapour react together according to the following equation.

$$2NO(g) + Br_2(g) \rightarrow 2NOBr(g) \qquad \Delta H^\ominus = -23\,kJ\,mol^{-1}$$

The reaction has an activation energy of $+5.4\,kJ\,mol^{-1}$.

What is the correct reaction pathway diagram for this reaction?

A



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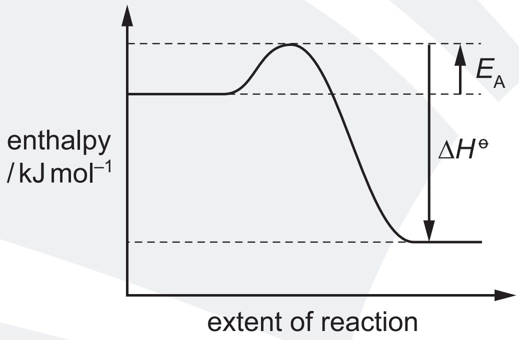
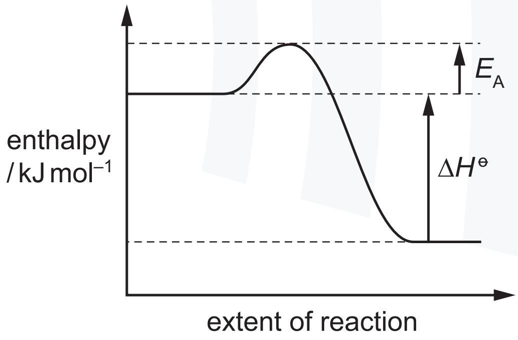
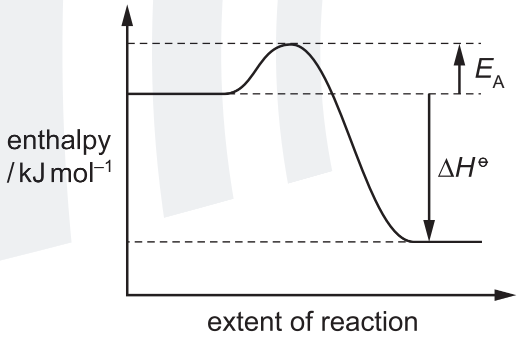


C



D



81 Which equation has an enthalpy change of reaction which corresponds to the standard enthalpy change of atomisation of chlorine?

A $\frac{1}{2}Cl_2(g) \rightarrow Cl(g)$

B $\frac{1}{2}Cl_2(l) \rightarrow Cl(g)$

C $Cl_2(g) \rightarrow 2Cl(g)$

D $Cl_2(l) \rightarrow 2Cl(g)$

82 Which equation shows the reaction that occurs during the standard enthalpy change of atomisation of bromine?

A $Br_2(l) \rightarrow 2Br(g)$

B $Br_2(g) \rightarrow 2Br(g)$

C $\frac{1}{2}Br_2(l) \rightarrow Br(g)$

D $\frac{1}{2}Br_2(g) \rightarrow Br(g)$

83 Gaseous phosphorus pentachloride can be decomposed into gaseous phosphorus trichloride and chlorine by heating. The table gives the bond energies.

bond	bond energy / $kJ\,mol^{-1}$
P–C*l* (in both chlorides)	330
C*l*–C*l*	242

What is the enthalpy change for the decomposition of PCl_5 to PCl_3 and Cl_2?

A $-418\,kJ\,mol^{-1}$

B $-88\,kJ\,mol^{-1}$

C $+88\,kJ\,mol^{-1}$

D $+418\,kJ\,mol^{-1}$

84 The equation for the complete combustion of propan-1-ol is shown.

$$CH_3CH_2CH_2OH(l) + 4\tfrac{1}{2}O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$$

Standard enthalpy changes of formation are given.

compound	$CH_3CH_2CH_2OH(l)$	$CO_2(g)$	$H_2O(l)$
$\Delta H_f^\ominus$	–303 kJ mol^{-1}	–394 kJ mol^{-1}	–286 kJ mol^{-1}

What is the standard enthalpy change of combustion of propan-1-ol, in kJ mol^{-1}?

A $-394 - 286 - 303$

B $303 - (4 \times 286) - (3 \times 394)$

C $394 + 286 - 303$

D $(3 \times 394) + (4 \times 286) + 303$

85 Enthalpy changes, ΔH, can be positive or negative.

Which row is correct?

	ΔH positive	ΔH negative
A	atomisation	bond breaking
B	bond breaking	neutralisation
C	bond making	combustion
D	combustion	bond making

86 Methanol may be prepared by the reaction between carbon monoxide and hydrogen.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$$

The relevant average bond energies are given below.

$E(C{\equiv}O)$	$1077\,kJ\,mol^{-1}$
$E(C-O)$	$360\,kJ\,mol^{-1}$
$E(C-H)$	$410\,kJ\,mol^{-1}$
$E(H-H)$	$436\,kJ\,mol^{-1}$
$E(O-H)$	$460\,kJ\,mol^{-1}$

What is the enthalpy change of this reaction?

A $-537\,kJ\,mol^{-1}$

B $-101\,kJ\,mol^{-1}$

C $+101\,kJ\,mol^{-1}$

D $+537\,kJ\,mol^{-1}$

87 The enthalpy change of reaction 1 is $-114\,kJ\,mol^{-1}$.

$$2NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l) \qquad \text{reaction 1}$$

By using this information, what is the most likely value for the enthalpy change of reaction 2?

$$Ba(OH)_2(aq) + 2HCl(aq) \rightarrow BaCl_2(aq) + 2H_2O(l) \qquad \text{reaction 2}$$

A $-57\,kJ\,mol^{-1}$ **B** $-76\,kJ\,mol^{-1}$ **C** $-114\,kJ\,mol^{-1}$ **D** $-228\,kJ\,mol^{-1}$

[S'18 2 Q7]

88 In a calorimetric experiment 1.60 g of a fuel are burnt. 45.0% of the energy released is absorbed by 200 g of water. The temperature of the water rises from 18.0 °C to 66.0 °C.

What is the total energy released per gram of fuel burnt (to 3 significant figures)?

A 25 100 J **B** 55 700 J **C** 89 200 J **D** 143 000 J

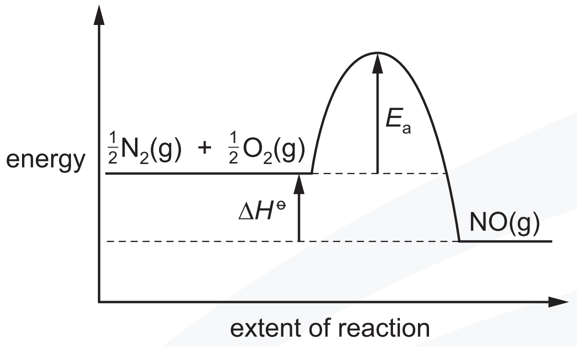
89 In the high temperatures of car engines, nitrogen reacts with oxygen to give nitrogen monoxide.

$$\tfrac{1}{2}N_2(g) + \tfrac{1}{2}O_2(g) \rightarrow NO(g) \qquad \Delta H^{\ominus} = +90\,\text{kJ mol}^{-1}$$

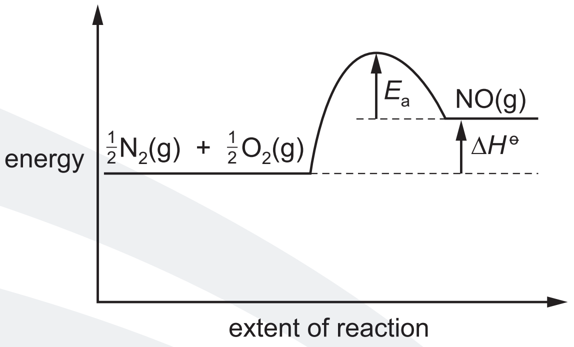
This reaction has activation energy E_a.

Which reaction pathway diagram could correctly represent this reaction?

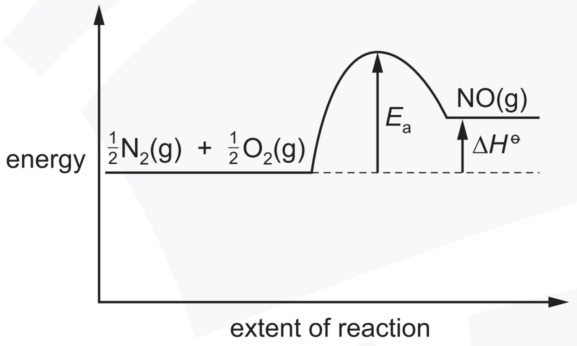
A



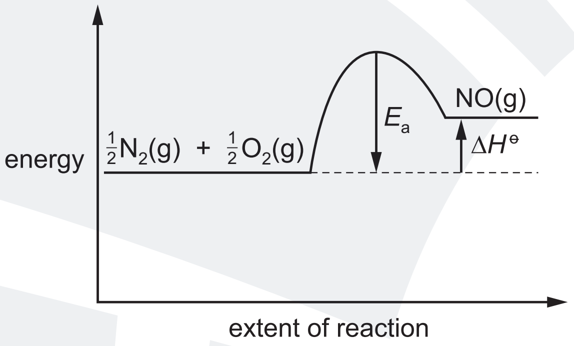
B



C



D



90 Which expression gives the standard enthalpy change of combustion of methane?

A $\Delta H_f^{\ominus}(CH_4) + \Delta H_f^{\ominus}(CO_2) - 2\Delta H_f^{\ominus}(H_2O)$

B $\Delta H_f^{\ominus}(CO_2) + 2\Delta H_f^{\ominus}(H_2O) + \Delta H_f^{\ominus}(CH_4)$

C $\Delta H_f^{\ominus}(CH_4) + 2\Delta H_f^{\ominus}(H_2O) - \Delta H_f^{\ominus}(CO_2)$

D $\Delta H_f^{\ominus}(CO_2) + 2\Delta H_f^{\ominus}(H_2O) - \Delta H_f^{\ominus}(CH_4)$

91 The following data are needed for this question.

$\Delta H_f^\ominus(CO(g)) = -111\,kJ\,mol^{-1}$

$\Delta H_f^\ominus(CO_2(g)) = -394\,kJ\,mol^{-1}$

$\Delta H_f^\ominus(Fe_2O_3(s)) = -822\,kJ\,mol^{-1}$

Carbon monoxide reacts with iron(III) oxide.

$$3CO(g) + Fe_2O_3(s) \rightarrow 3CO_2(g) + 2Fe(s)$$

What is the enthalpy change when 55.8 g of iron are produced by this reaction?

A –27.0 kJ **B** –13.5 kJ **C** +13.5 kJ **D** +27.0 kJ

92 All the reactants and products of an exothermic reaction are gaseous.

Which statement about this reaction is correct?

A The total bond energy of the products is less than the total bond energy of the reactants, and ΔH for the reaction is negative.

B The total bond energy of the products is less than the total bond energy of the reactants, and ΔH for the reaction is positive.

C The total bond energy of the products is more than the total bond energy of the reactants, and ΔH for the reaction is negative.

D The total bond energy of the products is more than the total bond energy of the reactants, and ΔH for the reaction is positive.

93 In calculating the enthalpy change, ΔH, of an experiment involving solutions, the mass of the solution, m, specific heat capacity of the solution, c, and the temperature change, ΔT, are needed.

$$\Delta T = T_{final} - T_{initial}$$

Which expression for ΔH is correct?

A $\Delta H = \dfrac{mc}{\Delta T}$ **B** $\Delta H = \dfrac{-mc}{\Delta T}$ **C** $\Delta H = mc\Delta T$ **D** $\Delta H = -mc\Delta T$

[W’17 2 Q6]

94 The following data are needed for this question.

$\Delta H_f^\ominus$ (P_4O_{10}(s)) = –3012 kJ mol^{-1}

$\Delta H_f^\ominus$ (H_2O(l)) = –286 kJ mol^{-1}

$\Delta H_f^\ominus$ (H_3PO_4(s)) = –1279 kJ mol^{-1}

What is $\Delta H^\ominus$ for the reaction shown?

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

A –9844 kJ mol^{-1}

B –388 kJ mol^{-1}

C –97 kJ mol^{-1}

D +2019 kJ mol^{-1}

95 Sulfur can be oxidised in two ways.

$$S(s) + O_2(g) \rightarrow SO_2(g) \qquad \Delta H^\ominus = -296.5\ \text{kJ mol}^{-1}$$

$$2S(s) + 3O_2(g) \rightarrow 2SO_3(g) \qquad \Delta H^\ominus = -791.4\ \text{kJ mol}^{-1}$$

Sulfur trioxide can be made from sulfur dioxide and oxygen.

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

What is the standard enthalpy change for this reaction?

A –1384.4 kJ mol^{-1}

B –989.8 kJ mol^{-1}

C –494.9 kJ mol^{-1}

D –198.4 kJ mol^{-1}

[W'17 3 Q8]

96 200 g of water are at 25 °C.

The water is heated to 75 °C by burning 2 g of ethanol.

What is the amount of energy transferred to the water?

A 0.418 kJ **B** 10.4 kJ **C** 41.8 kJ **D** 62.7 kJ

[W'17 3 Q11]

97 Which equation represents the standard enthalpy change of formation of water?

A $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(g)$

B $H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$

C $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$

D $2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$

[M'18 2 Q8]

98 Hess' Law and bond energy data can be used to calculate the enthalpy change of a reaction.

Bromoethane, CH_3CH_2Br, can be made by reacting ethene with hydrogen bromide.

$$CH_2{=}CH_2 + HBr \rightarrow CH_3CH_2Br$$

What is the enthalpy change for this reaction?

A −674 kJ mol^{-1}

B −64 kJ mol^{-1}

C +186 kJ mol^{-1}

D +346 kJ mol^{-1}

[M'18 2 Q9]

99 Enthalpy changes of combustion can be used to determine enthalpy changes of formation. The following equation represents the enthalpy change of formation of butane.

$$4C(s) + 5H_2(g) \rightarrow C_4H_{10}(g)$$

By using the following standard enthalpy of combustion data, what is the value of the standard enthalpy change of formation, $\Delta H_f^{\ominus}$, of butane?

substance	$\Delta H_c^{\ominus}$ / kJ mol^{-1}
C(s)	–394
$H_2(g)$	–286
$C_4H_{10}(g)$	–2877

A –5883 kJ mol^{-1}

B –129 kJ mol^{-1}

C +129 kJ mol^{-1}

D +2197 kJ mol^{-1}

[S'18 1 Q7]

SECTION B

For the following questions, the responses A to D should be selected on the basis of

A	B	C	D
1, 2 and **3** are correct	**1** and **2** only are correct	**2** and **3** only are correct	**1** only is correct

1 The conversion of graphite has only a small positive value of ΔH.

$$\text{C (graphite)} \rightarrow \text{C (diamond)} \qquad \Delta H = +2.1\ \text{kJ mol}^{-1}$$

However, the production of synthetic diamonds using this reaction is very difficult.

Which statements help to explain this?

1 The activation energy of the reaction is large.

2 An equilibrium exists between diamond and graphite.

3 Only exothermic reactions can be made to occur readily.

2 Sodium ions can be formed from sodium atoms.

$$\text{Na(s)} \rightarrow \text{Na}^{+}\text{(g)}$$

Which quantities are required to calculate the enthalpy change of formation of gaseous sodium ions?

1 enthalpy change of atomisation of sodium

2 first ionisation energy of sodium

3 enthalpy change of formation of sodium

3 The conversion of graphite into diamond is an endothermic reaction ($\Delta H = +3\ \text{kJ mol}^{-1}$).

$$\text{C(graphite)} \rightarrow \text{C(diamond)}$$

Which statements are correct?

1 The enthalpy change of atomisation of diamond is smaller than that of graphite.

2 The bond energy of the C–C bonds in graphite is greater than that in diamond.

3 The enthalpy change of combustion of diamond is greater than that of graphite.

4 The diagram illustrates the energy changes of a set of reactions.

$$\begin{tikzcd}
R \arrow[r, "\Delta H = -134\ \text{kJ mol}^{-1}"] \arrow[d] & S \arrow[d, "\Delta H = +92\ \text{J mol}^{-1}"] \\
T \arrow[r, "\Delta H = -75\ \text{kJ mol}^{-1}"] & U
\end{tikzcd}$$

Which of the following statements are correct?

1 The enthalpy change for the transformation **U** $\longrightarrow$ **R** is $+42$ kJ mol^{-1} .

2 The enthalpy change for the transformation **T** $\longrightarrow$ **S** is endothermic.

3 The enthalpy change for the transformation **R** $\longrightarrow$ **T** is -33 kJ mol^{-1} .

5 Which of the enthalpy changes of the following reactions can **only** be obtained by application of Hess' Law?

1 The hydration of anhydrous copper sulphate to form crystals of $CuSO_4.5H_2O$.

2 The formation of methane from its elements.

3 The combustion of glucose, $C_6H_{12}O_6$.

6 Nitrogen and oxygen react in a hot car engine to form nitrogen monoxide which is a serious pollutant in our cities and in the countryside. However, nitrogen and oxygen do not react at room temperature.

Which statements help to explain why nitrogen and oxygen do not react at room temperature?

1 The reaction is endothermic.

2 A high activation energy is required.

3 Nitrogen has a high bond energy.

7 Hydrazine, N_2H_4, is widely used as a rocket fuel because it reacts with oxygen as shown, producing 'environmentally friendly' gases.

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g) \qquad \Delta H = -534\ \text{kJ mol}^{-1}$$

Despite its use as a rocket fuel, hydrazine does not spontaneously burn in oxygen.

Why does hydrazine not burn spontaneously?

1 The activation energy is too high.

2 The N≡N bond is very strong.

3 Hydrazine is a liquid.

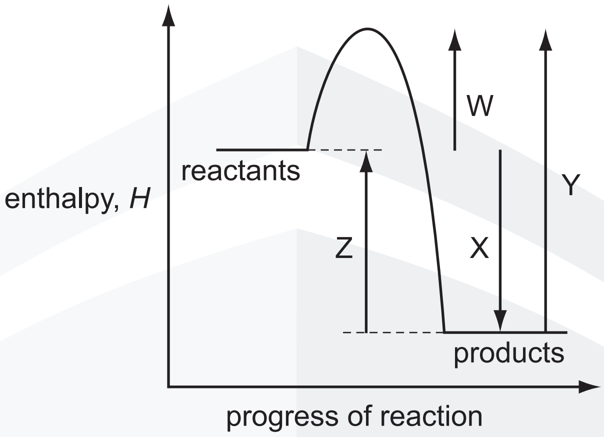
8 For which reactions does the value of $\Delta H^{\ominus}$ represent **both** a standard enthalpy change of combustion **and** a standard enthalpy change of formation?

1 $C(s) + O_2(g) \rightarrow CO_2(g)$

2 $2C(s) + O_2(g) \rightarrow 2CO(g)$

3 $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

9 An energy profile diagram is shown.



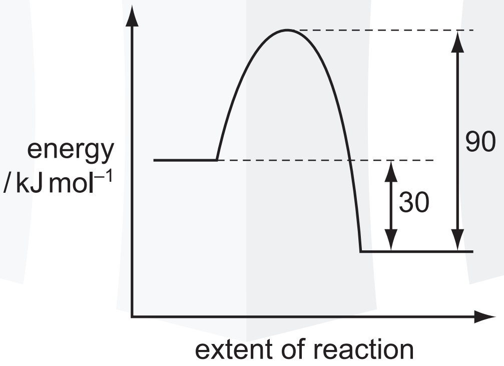
What do the labels on the diagram represent?

1 W = ΔH of the forward reaction, Y = E_a of the backward reaction

2 Z = ΔH of the backward reaction, Y = E_a of the backward reaction

3 X = ΔH of the forward reaction, W = E_a of the forward reaction

10 The diagram shows the reaction pathway for a reversible reaction.



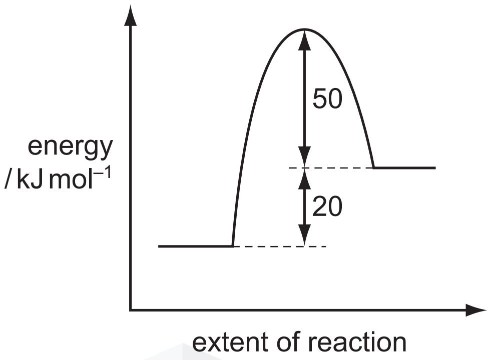
Which statements are correct?

1 The enthalpy change for the backward reaction is $+90\,kJ\,mol^{-1}$.

2 The forward reaction is exothermic.

3 The enthalpy change for the forward reaction is $-30\,kJ\,mol^{-1}$.

11 The reaction pathway for a reversible reaction is shown below.



Which statements are correct?

1 The enthalpy change for the backward reaction is –20 kJ mol^{-1}.

2 The forward reaction is endothermic.

3 The activation energy for the forward reaction is +70 kJ mol^{-1}.

12 The bond energy of the Br–O bond is 235 kJ mol^{-1}.

Which reactions are exothermic?

1 $OH\bullet + HBr \rightarrow H_2 + BrO\bullet$

2 $OH\bullet + HBr \rightarrow H_2O + Br\bullet$

3 $H\bullet + HBr \rightarrow H_2 + Br\bullet$

13 Which names can be applied to the enthalpy change of the reaction shown?

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l)$$

1 enthalpy change of formation

2 enthalpy change of combustion

3 enthalpy change of hydration

14 Compared with the HI molecule, the bond ……P…… of the HBr molecule is ……Q……

Which pairs of words correctly complete the above sentence?

	P	Q
1	energy	greater
2	length	less
3	polarity	greater

15 Which statements are correct for **all** exothermic reactions?

1 ΔH for the reaction is negative.

2 On a reaction pathway diagram the products are shown lower than the reactants.

3 The reaction will happen spontaneously.

16 The diagram illustrates the enthalpy changes of a set of reactions

$$\begin{tikzcd}
R \arrow[r, "\Delta H = -134\,\mathrm{kJ\,mol^{-1}}"] \arrow[d] & S \arrow[d, "\Delta H = +92\,\mathrm{kJ\,mol^{-1}}"] \\
T \arrow[r, "\Delta H = -75\,\mathrm{kJ\,mol^{-1}}"] & U
\end{tikzcd}$$

Which statements are correct?

1 The enthalpy change for the transformation U → R is $+42\,\mathrm{kJ\,mol^{-1}}$.

2 The enthalpy change for the transformation T → S is endothermic.

3 The enthalpy change for the transformation R → T is $-33\,\mathrm{kJ\,mol^{-1}}$.

17 A chemist puts a sample of dilute aqueous hydrochloric acid into beaker 1. She adds a sample of zinc and measures the rate of production of hydrogen gas.

She then puts a different sample of dilute aqueous hydrochloric acid into beaker 2. She adds a different sample of zinc and measures the rate of production of hydrogen gas.

The rate of the reaction in beaker 2 is greater than the rate of the reaction in beaker 1.

Which factors **could** help to explain this observation?

1 The reaction in beaker 1 has a higher activation energy than the reaction in beaker 2.

2 The zinc in beaker 1 is in larger pieces than the zinc in beaker 2.

3 The acid in beaker 1 is at a lower concentration than the acid in beaker 2.

18 An ethanol burner can be used to heat water. If appropriate measurements are taken, a value for the enthalpy of combustion of ethanol can be calculated. The equation

$$\text{heat transferred} = -mc\Delta T$$

is used as part of the calculation.

Which symbols are correctly described?

1 ΔT is the change in temperature of the water.

2 *m* is the mass of water used in the experiment.

3 *c* is the specific heat capacity of ethanol.

19 Calcium reacts with water to form calcium hydroxide and hydrogen.

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

The standard enthalpy change for this reaction is $-414\,\text{kJ mol}^{-1}$.

What further information is needed in order to calculate the standard enthalpy change of formation of calcium hydroxide, ΔH_f° $Ca(OH)_2(s)$?

1 ΔH_f° for $H_2O(l)$

2 ΔH_f° for $H_2(g)$

3 first and second ionisation energies of Ca

[S'18 1 Q33]

20 For which reactions does the value of ΔH° represent **both** a standard enthalpy change of combustion **and** a standard enthalpy change of formation?

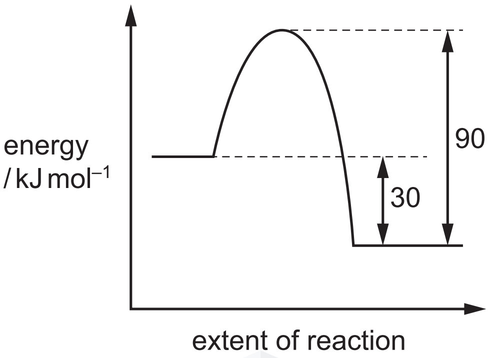
1 $C(s) + O_2(g) \rightarrow CO_2(g)$

2 $2C(s) + O_2(g) \rightarrow 2CO(g)$

3 $CO(g) + \frac{1}{2}O_2(g) \rightarrow CO_2(g)$

[S'18 2 Q33]

21 The diagram shows the reaction pathway for a reversible reaction.



Which statements are correct?

1 The forward reaction is exothermic.

2 The enthalpy change for the forward reaction is –30 kJ mol^{-1}.

3 The enthalpy change for the backward reaction is +90 kJ mol^{-1}.

[S'18 3 Q34]

ENERGETICS WS 2

1 The unsaturated hydrocarbon **Z** is obtained by cracking hexane and is important in the chemical industry.

The standard enthalpy change of combustion of **Z** is –2059 kJ mol^{-1}.

(d) Define the term *standard enthalpy change of combustion*.

....

.... [2]

When 0.47 g of **Z** were completely burnt in air, the heat produced raised the temperature of 200 g of water by 27.5 °C.

(e) **(i)** Calculate the amount of heat released in this experiment.

(ii) Use the data above and your answer to **(i)** to calculate the relative molecular mass of **Z**.

[4]

2 **(i)** Explain how enthalpy changes, ΔH values, for covalent bonded molecules can be calculated from bond energies.

....

....

(ii) Use bond energies from the *Data Booklet* to calculate ΔH for the following dissociation.

$$2HI(g) \rightarrow H_2(g) + I_2(g)$$

[3]

3 **(d)** The equation for the complete combustion of ethyne is given below.
Use appropriate bond energy data from the *Data Booklet* to calculate a value for the enthalpy change of combustion of ethyne.

$$C_2H_2(g) + \frac{5}{2}O_2(g) \longrightarrow 2CO_2(g) + H_2O(g)$$

[3]

(e) The value for the standard enthalpy change of combustion of ethyne is $-1300\,kJ\,mol^{-1}$.

(i) Define the term *standard enthalpy change of combustion*.

....

....

....

(ii) Explain why your answer to **(d)** does not have the same value as the standard enthalpy change of combustion.

....

....

[3]

4 Urea, $CO(NH_2)_2$, is a naturally occurring substance which can be hydrolysed with water to form ammonia according to the following equation.

$$H_2O(l) + CO(NH_2)_2(aq) \longrightarrow CO_2(aq) + 2NH_3(aq)$$

The standard enthalpy changes of formation of water, urea, carbon dioxide and ammonia (in aqueous solution) are given below.

compound	$\Delta H_f^{\ominus}$/kJ mol^{-1}
$H_2O(l)$	−287.0
$CO(NH_2)_2(aq)$	−320.5
$CO_2(aq)$	−414.5
$NH_3(aq)$	−81.0

Use these data to calculate the standard enthalpy change for the hydrolysis of urea.

[2]

5 Alkanes such as methane, CH_4, undergo few chemical reactions. Methane will, however, react with chlorine but not with iodine.

Relevant standard enthalpy changes of formation for the reaction of methane with chlorine to form chloromethane, CH_3Cl, are given below.

	$\Delta H_f^\ominus$/kJ mol^{-1}
CH_4	–75
CH_3Cl	–82
HCl	–92

(a) **(i)** Use the data to calculate $\Delta H^\ominus_{reaction}$ for the formation of CH_3Cl.

$$CH_4 + Cl_2 \rightarrow CH_3Cl + HCl$$

(ii) The corresponding reaction with iodine does **not** take place.

Use bond energy data from the *Data Booklet* to calculate a 'theoretical value' for $\Delta H_{reaction}$ for the following equation.

$$CH_4 + I_2 \rightarrow CH_3I + HI$$

(iii) Suggest why this reaction does **not** in fact occur.

....

....

[5]

6 Hydrazine reacts with oxygen according to the following equation.

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(g)$$

(i) Use the data in the table to calculate the standard enthalpy change of this reaction.

compound	$\Delta H_f^{\ominus}$/kJ mol^{-1}
$N_2H_4(l)$	50.6
$H_2O(g)$	–241.8

$\Delta H^{\ominus}$ = ... kJ mol^{-1}

(ii) Although the above reaction is highly exothermic, hydrazine does not burn spontaneously in oxygen.
Suggest a reason for this.

....

....

(iii) Suggest why using hydrazine as a rocket fuel could be regarded as being ‘environmentally friendly’.

....

....

[4]

7 The standard enthalpy change of combustion of C_2H_2, $\Delta H^{\ominus}_{c}$, is $-1300\,\text{kJ mol}^{-1}$ at 298 K.

Values of relevant standard enthalpy changes of formation, $\Delta H^{\ominus}_{f}$, measured at 298 K, are given in the table.

substance	$\Delta H^{\ominus}_{f}$ / kJ mol^{-1}
$CO_2(g)$	–394
$H_2O(l)$	–286

(i) Write balanced equations, with state symbols, that represent

the standard enthalpy change of combustion, $\Delta H^{\ominus}_{c}$, of C_2H_2, and

....

the standard enthalpy change of formation, $\Delta H^{\ominus}_{f}$, of C_2H_2.

....

(ii) Use the data above and your answer to **(i)** to calculate the standard enthalpy change of formation, $\Delta H^{\ominus}_{f}$, of C_2H_2.
Show clearly whether the standard enthalpy change of formation of C_2H_2 has a positive or negative value.

[6]

8 Ethanol, C_2H_5OH, is a most important industrial chemical and is used as a solvent, a fuel and an intermediate in large scale organic synthesis.

Ethanol is prepared industrially by the reaction of ethene and steam in the presence of a catalyst.

$$C_2H_4(g) + H_2O(g) \longrightarrow C_2H_5OH(g)$$

The standard enthalpy change of the reaction can be determined by using the standard enthalpy changes of combustion, $\Delta H_c^\ominus$, at 298 K.

	$\Delta H_c^\ominus$ / kJ mol^{-1}
$C_2H_4(g)$	–1411
$C_2H_5OH(l)$	–1367

(a) Calculate the standard enthalpy change for the following reaction.

$$C_2H_4(g) + H_2O(l) \longrightarrow C_2H_5OH(l)$$

[2]

(b) (i) Define the term *standard enthalpy change of combustion*.

....

....

....

(ii) Explain why the state symbols for water and ethanol given in the equation in **(a)** have been changed from those quoted in the industrial process.

....

....

....

(iii) Write the equation for the complete combustion of ethanol.

....

[4]

9 Carbon, hydrogen and ethene each burn exothermically in an excess of air.

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta H^{\ominus}_c = -393.7\,kJ\,mol^{-1}$$

$$H_2(g) + ½O_2(g) \rightarrow H_2O(l) \qquad \Delta H^{\ominus}_c = -285.9\,kJ\,mol^{-1}$$

$$C_2H_4(g) + 3O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l) \qquad \Delta H^{\ominus}_c = -1411.0\,kJ\,mol^{-1}$$

Use the data to calculate the standard enthalpy change of formation, $\Delta H^{\ominus}_f$, in kJ mol^{-1}, of ethene at 298 K.

$$2C(s) + 2H_2(g) \rightarrow C_2H_4(g)$$

$\Delta H^{\ominus}_f$ = ... kJ mol^{-1} [3]

10 Calculate the standard enthalpy change of formation of CS_2 from the following data. Include a sign in your answer.

standard enthalpy change of combustion of CS_2 = $-1110\,kJ\,mol^{-1}$

standard enthalpy change of formation of CO_2 = $-395\,kJ\,mol^{-1}$

standard enthalpy change of formation of SO_2 = $-298\,kJ\,mol^{-1}$

[3]

11 Ketene, C_2H_2O, is a member of a class of unsaturated organic compounds that is widely used in pharmaceutical research for the synthesis of organic compounds.

Use the data below to calculate the standard enthalpy change of formation of ketene.

	$\Delta H^{\ominus}$/kJ mol^{-1}
standard enthalpy change of formation of CO_2	–395
standard enthalpy change of combustion of H_2	–286
standard enthalpy change of combustion of $CH_2{=}C{=}O$	–1028

[6]

12 **(a)** **(i)** What is meant by the *standard enthalpy change of formation*, $\Delta H^{\ominus}_f$, of a compound? Explain what is meant by the term *standard*.

....

....

....

(ii) Write an equation, with state symbols, for the $\Delta H^{\ominus}_f$ of water.

....

(iii) Explain why the $\Delta H^{\ominus}_f$ for water is identical to the standard enthalpy change of combustion of hydrogen.

....

....[4]

(b) When calcium is placed in water, aqueous calcium hydroxide is formed and hydrogen is given off.

(i) Write the equation for the reaction of calcium with water.

....

(ii) When 1.00 g of calcium is placed in 200 g of water, the temperature increases by 12.2 °C when the reaction is completed. The specific heat capacity of water, *c*, is 4.2 J g^{-1} K^{-1}.

Calculate the heat released in the experiment.

(iii) Calculate the standard enthalpy change of reaction in k J mol^{-1} for your equation in **(b)(i)**.

[4]

(c) **(i)** State *Hess' Law*.

....

....

....

(ii) Use Hess' Law and your result in **(b)(iii)** to calculate the $\Delta H^{\ominus}_f$ of $Ca(OH)_2$ (aq). You also need the $\Delta H^{\ominus}_f$ of water which is –286 k J mol^{-1}.

[4]

(d) Calculate the volume of hydrogen, measured at room temperature and pressure, liberated in the experiment described in **(b)(ii)**.

[2]

13 For some chemical reactions, such as the thermal decomposition of potassium hydrogencarbonate, $KHCO_3$, the enthalpy change of reaction cannot be measured directly.

In such cases, the use of Hess' Law enables the enthalpy change of reaction to be calculated from the enthalpy changes of other reactions.

(a) State Hess' Law.

....

....

.... [2]

In order to determine the enthalpy change for the thermal decomposition of potassium hydrogencarbonate, two separate experiments were carried out.

experiment 1

30.0 cm^3 of 2.00 mol dm^{-3} hydrochloric acid (an excess) was placed in a conical flask and the temperature recorded as 21.0 °C.
When 0.0200 mol of potassium carbonate, K_2CO_3, was added to the acid and the mixture stirred with a thermometer, the maximum temperature recorded was 26.2 °C.

(b) **(i)** Construct a balanced equation for this reaction.

....

(ii) Calculate the quantity of heat produced in **experiment 1**, stating your units.
Use relevant data from the *Data Booklet* and assume that all solutions have the same specific heat capacity as water.

(iii) Use your answer to **(ii)** to calculate the enthalpy change per mole of K_2CO_3.
Give your answer in kJ mol^{-1} and include a sign in your answer.

(iv) Explain why the hydrochloric acid must be in an excess.

....

.... [4]

experiment 2

The experiment was repeated with 0.0200 mol of potassium hydrogencarbonate, $KHCO_3$.
All other conditions were the same.
In the second experiment, the temperature fell from 21.0 °C to 17.3 °C.

(c) **(i)** Construct a balanced equation for this reaction.

....

(ii) Calculate the quantity of heat absorbed in **experiment 2**.

(iii) Use your answer to **(ii)** to calculate the enthalpy change per mole of $KHCO_3$.
Give your answer in kJ mol^{-1} and include a sign in your answer.

[3]

(d) When $KHCO_3$ is heated, it decomposes into K_2CO_3, CO_2 and H_2O.

$$2KHCO_3 \rightarrow K_2CO_3 + CO_2 + H_2O$$

Use Hess' Law and your answers to **(b)(iii)** and **(c)(iii)** to calculate the enthalpy change for this reaction.
Give your answer in kJ mol^{-1} and include a sign in your answer.

[2]

14 Methanol, CH_3OH, is considered to be a possible alternative to fossil fuels, particularly for use in vehicles.

Methanol can be produced from fossil fuels and from agricultural waste. It can also be synthesised from carbon dioxide and hydrogen.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of formation of carbon dioxide.

equation ....

definition ....

....

.... [3]

(b) Relevant $\Delta H_f^{\ominus}$ values for the reaction that synthesises methanol are given in the table.

compound	$\Delta H_f^{\ominus}$ / kJ mol^{-1}
$CO_2(g)$	–394
$CH_3OH(g)$	–201
$H_2O(g)$	–242

(i) Use these values to calculate $\Delta H^{\ominus}_{reaction}$ for this synthesis of methanol.

Include a sign in your answer.

$$CO_2(g) + 3H_2(g) \rightleftharpoons CH_3OH(g) + H_2O(g)$$

$\Delta H^{\ominus}_{reaction}$ = ... kJ mol^{-1}

(ii) Suggest **one** possible environmental advantage of this reaction. Explain your answer.

....

....

[5]

15 Alcohols such as methanol, CH_3OH, are considered to be possible replacements for fossil fuels because they can be used in car engines.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of combustion, $\Delta H^{\ominus}_{c}$, for methanol at 298 K.

equation ....

definition ....

....

.... [3]

Methanol may be synthesised from carbon monoxide and hydrogen. Relevant $\Delta H^{\ominus}_{c}$ values for this reaction are given in the table below.

compound	$\Delta H^{\ominus}_{c}$ / kJ mol^{-1}
CO(g)	–283
H_2(g)	–286
CH_3OH(g)	–726

(b) Use these values to calculate $\Delta H^{\ominus}_{reaction}$ for the synthesis of methanol, using the following equation. Include a sign in your answer.

$$CO(g) + 2H_2(g) \rightarrow CH_3OH(g)$$

$\Delta H^{\ominus}_{reaction}$ = ...kJ mol^{-1}

[3]

16 With the prospect that fossil fuels will become increasingly scarce in the future, many compounds are being considered for use in internal combustion engines. One of these is DME or dimethyl ether, CH_3OCH_3. DME is a gas which can be synthesised from methanol. Methanol can be obtained from biomass, such as plant waste from agriculture.

(a) Define, with the aid of an equation which includes state symbols, the standard enthalpy change of combustion, $\Delta H^{\ominus}_{c}$, for DME at 298 K.

equation ....

definition ....

....

.... [3]

(b) DME may be synthesised from methanol. Relevant enthalpy changes of formation, $\Delta H^{\ominus}_{f}$, for this reaction are given in the table below.

compound	$\Delta H^{\ominus}_{f}$ / kJ mol^{-1}
$CH_3OH(l)$	–239
$CH_3OCH_3(g)$	–184
$H_2O(l)$	–286

Use these values to calculate $\Delta H^{\ominus}_{reaction}$ for the synthesis of DME, using the following equation. Include a sign in your answer.

$$2CH_3OH(l) \rightarrow CH_3OCH_3(g) + H_2O(l)$$

$\Delta H^{\ominus}_{reaction}$ = ... kJ mol^{-1}
[3]

17 The standard enthalpy changes of formation of $NH_3(g)$ and $H_2O(g)$ are as follows.

$NH_3(g)$, $\Delta H_f^{\ominus} = -46.0\,\text{kJ mol}^{-1}$ $H_2O(g)$, $\Delta H_f^{\ominus} = -242\,\text{kJ mol}^{-1}$

Use these data and the value of $\Delta H^{\ominus}_{\text{reaction}}$ given below to calculate the standard enthalpy change of formation of NO(g).
Include a sign in your answer.

$$4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g) \qquad \Delta H^{\ominus} = -906\,\text{kJ mol}^{-1}$$

[4]

18 A 1.00 cm^3 sample of $C_{14}H_{30}$ was completely burnt in air.
The heat produced raised the temperature of 250 g of water by 34.6 °C.
Assume no heat losses occurred during this experiment.
The density of $C_{14}H_{30}$ is 0.763 g cm^{-3}.

(i) Use relevant data from the *Data Booklet* to calculate the amount of heat released in this experiment.

(ii) Use the data above and your answer to **(i)** to calculate the energy produced by the combustion of 1 mol of $C_{14}H_{30}$.

[5]

19 Ethanal reacts with hydrogen cyanide, in the presence of a small amount of NaCN, as shown.

$$CH_3CHO + HCN \rightarrow CH_3CH(OH)CN$$

(a) Use bond energies from the *Data Booklet* to calculate the enthalpy change for this reaction. Include a sign with your answer.

enthalpy change = ... kJ mol^{-1} [3]

20 Chemical reactions are accompanied by enthalpy changes.

(a) Explain the meaning of the term *standard enthalpy change* of reaction.

....

....

.... [2]

(b) The enthalpy change of hydration of anhydrous magnesium sulfate, ΔH_{hyd} $MgSO_4$, can be calculated by carrying out two separate experiments.

In the first experiment 45.00 g of water was weighed into a polystyrene cup and 3.01 g of $MgSO_4$ was added and stirred until it was completely dissolved. The temperature of the water rose from 23.4 °C to 34.7 °C.

(i) Calculate the amount of heat energy transferred to the water during this dissolving process.

You can assume that the specific heat capacity of the solution is the same as that of water, $4.18\,J\,g^{-1}\,K^{-1}$.

heat energy = ... J [1]

(ii) Calculate the amount, in moles, of $MgSO_4$ dissolved.

amount = ... mol [1]

(iii) Calculate the enthalpy change of solution, ΔH_{soln}, of $MgSO_4$(s).

You must include a sign with your answer.

ΔH_{soln}, of $MgSO_4$(s) = ... kJ mol^{-1} [1]

In the second experiment, the enthalpy change of solution for the hydrated salt, $MgSO_4.7H_2O$(s), was calculated and found to be +9.60 kJ mol^{-1}.

(iv) Use the equation below for the hydration of anhydrous magnesium sulfate to construct a suitable, fully labelled energy cycle that will allow you to calculate the enthalpy change for this reaction, ΔH_{hyd} $MgSO_4$.

$$MgSO_4(s) + 7H_2O(l) \rightarrow MgSO_4.7H_2O(s)$$

[1]

(v) Calculate the enthalpy change for this reaction, ΔH_{hyd} $MgSO_4$. Include a sign in your answer.

ΔH_{hyd} $MgSO_4$ = ... kJ mol^{-1} [1]

21 **a** A 1.00 cm^3 sample of $C_{14}H_{30}$ was completely burnt in air.
The heat produced raised the temperature of 250 g of water by 34.6 °C.
Assume no heat losses occurred during this experiment.
The density of $C_{14}H_{30}$ is 0.763 g cm^{-3}.

(i) Use relevant data from the *Data Booklet* to calculate the amount of heat released in this experiment.

(ii) Use the data above and your answer to **(i)** to calculate the energy produced by the combustion of 1 mol of $C_{14}H_{30}$.

[5]

22 The elements in Group 17, the halogens, and their compounds, show many similarities and trends in their properties. Some data are given for the elements fluorine to iodine.

element	bond energy / kJ mol^{-1}	standard enthalpy change of atomisation, $\Delta H^{\ominus}_{at}$ / kJ mol^{-1}	boiling point of element / K	boiling point of **hydrogen halide** / K
fluorine, F–F	158	79	85	293
chlorine, C*l*–C*l*	242	121	238	188
bromine, Br–Br	193	112	332	206
iodine, I–I	151	107	457	238

(a) (i) Explain the meaning of the term *standard enthalpy change of atomisation*.

....

....

.... [3]

(ii) For fluorine and chlorine, the enthalpy changes of atomisation are half the value of the bond energies.

For bromine and iodine, the enthalpy changes of atomisation are much more than half the value of the bond energies.

Suggest a reason for this difference.

....

....

.... [1]

(iii) The standard enthalpy of formation of iodine monochloride, IC*l*, is –24.0 kJ mol^{-1}.

Use this information and the bond energies of iodine and chlorine to calculate the I–C*l* bond energy.

I–C*l* bond energy = ... kJ mol^{-1} [2]

23 For many compounds the enthalpy change of formation cannot be calculated directly. An indirect method based on enthalpy changes of combustion can be used.

The enthalpy change of combustion can be found by a calorimetry experiment in which the heat energy given off during combustion is used to heat a known mass of water and the temperature change recorded.

(a) **(i)** Explain the meaning of the term *standard enthalpy change of combustion*.

....

....

.... [3]

(ii) Write the equation for the complete combustion of ethanol, C_2H_5OH.

.... [1]

(b) In an experiment to determine the enthalpy change of combustion of ethanol, 0.23 g of ethanol was burned and the heat given off raised the temperature of 100 g of water by 16.3 °C.

(i) Calculate the heat energy change, *q*, during the combustion of 0.23 g of ethanol.

q = ... J [1]

(ii) Calculate the enthalpy change on burning 1 mole of ethanol. Include a sign in your answer.

ΔH = ... kJ mol^{-1} [1]

(iii) Suggest **two** reasons why the value for the enthalpy change of combustion of ethanol determined by a simple laboratory calorimetry experiment is likely to be lower than the true value.

....

....

....

.... [2]

The table gives some enthalpy change of combustion values.

substance	enthalpy change of combustion / kJ mol^{-1}
C(s)	–393.5
H_2(g)	–285.8
C_3H_7OH(l)	–2021.0

(i) Construct a labelled energy cycle to show how these values could be used to calculate the enthalpy change of formation of $C_3H_7OH(l)$, ΔH_f.

$$3C(s) + 4H_2(g) + \tfrac{1}{2}O_2(g) \xrightarrow{\Delta H_f} C_3H_7OH(l)$$

[3]

(ii) Calculate the enthalpy change of formation, ΔH_f, of $C_3H_7OH(l)$.

ΔH_f = ... kJ mol^{-1} [2]

24 For many compounds the enthalpy change of formation cannot be calculated directly. An indirect method based on enthalpy changes of combustion can be used.

The enthalpy change of combustion can be found by a calorimetry experiment in which the heat energy given off during combustion is used to heat a known mass of water and the temperature change recorded.

(a) **(i)** Explain the meaning of the term *standard enthalpy change of combustion*.

....

....

.... [3]

(ii) Write the equation for the complete combustion of ethanol, C_2H_5OH.

.... [1]

(b) In an experiment to determine the enthalpy change of combustion of ethanol, 0.23 g of ethanol was burned and the heat given off raised the temperature of 100 g of water by 16.3 °C.

(i) Calculate the heat energy change, q, during the combustion of 0.23 g of ethanol.

q = ... J [1]

(ii) Calculate the enthalpy change on burning 1 mole of ethanol. Include a sign in your answer.

ΔH = ... kJ mol^{-1} [1]

(iii) Suggest **two** reasons why the value for the enthalpy change of combustion of ethanol determined by a simple laboratory calorimetry experiment is likely to be lower than the true value.

....

....

....

.... [2]

(c) The table gives some enthalpy change of combustion values.

substance	enthalpy change of combustion / kJ mol^{-1}
$C(s)$	–393.5
$H_2(g)$	–285.8
$C_3H_7OH(l)$	–2021.0

(i) Construct a labelled energy cycle to show how these values could be used to calculate the enthalpy change of formation of $C_3H_7OH(l)$, ΔH_f.

$$3C(s) + 4H_2(g) + \frac{1}{2}O_2(g) \xrightarrow{\Delta H_f} C_3H_7OH(l)$$

[3]

(ii) Calculate the enthalpy change of formation, ΔH_f, of $C_3H_7OH(l)$.

ΔH_f = ... kJ mol^{-1} [2]

NOTES

Redox

This topic illustrates the relationship between electricity and chemical changes. Chemical reactions can be investigated by looking at electrode potentials.

a calculate oxidation numbers of elements in compounds and ions ☐

b describe and explain redox processes in terms of electron transfer and changes in oxidation number ☐

c use changes in oxidation numbers to help balance chemical equations ☐

6.1 Redox processes: electron transfer and changes in oxidation number (oxidation state)	a) calculate oxidation numbers of elements in compounds and ions b) describe and explain redox processes in terms of electron transfer and changes in oxidation number c) use changes in oxidation numbers to help balance chemical equations

REDOX

6.1 Redox processes: electron transfer and changes in oxidation number (oxidation state)	a) calculate oxidation numbers of elements in compounds and ions b) describe and explain redox processes in terms of electron transfer and changes in oxidation number c) use changes in oxidation numbers to help balance chemical equations

WHAT YOU NEED TO KNOW

Calculate oxidation numbers of elements in compounds and ions

Describe and explain redox processes in terms of electron transfer and changes in oxidation numbers

Use changes in oxidation numbers to help balance chemical equations

REDOX

A redox reaction is a reaction in which electron transfer takes place.

Redox reactions form one of the largest groups in which we classify chemical changes.

In a redox reactions, oxidation and reduction will **always** occur together.

REDOX DEFINITIONS

OXIDATION IS:	REDUCTION IS:
gain of oxygen	gain of hydrogen
loss of hydrogen	loss of oxygen
loss of electrons	gain of electrons
increase in oxidation state	decrease in oxidation state

OXIDATION

1. Oxidation of an element:

$$S(s) + O_2(g) \longrightarrow SO_2(g)$$

2. Oxidation of a compound:

$$CH_4(g) + 2O_2(g) \longrightarrow CO_2(g) + 2H_2O(l)$$

3. Oxidation using an indirect source of oxygen:

$$8KClO_3(s) + C_{12}H_{22}O_{11}(s) \longrightarrow 12CO_2(g) + 11H_2O(l) + 8KCl(s)$$

SKILL CHECK

In each of the following equations, state which reactant has been oxidised:

(a) $PbO + H_2 \longrightarrow Pb + H_2O$

(b) $CO + Ag_2O \longrightarrow 2Ag + CO_2$

(c) $2Mg + CO_2 \longrightarrow 2MgO + C$

REDUCTION

1. Reduction to an element:

$$2HgO(s) \longrightarrow 2Hg(l) + O_2(g)$$

2. Reduction of one element by another:

$$2CuO(s) + C(s) \longrightarrow 2Cu(s) + CO_2(g)$$

SKILL CHECK

In each of the following equations, state which reactant has been reduced:

(a) $5CO + I_2O_2 \longrightarrow 5CO_2 + I_2$

(b) $2H_2S + SO_2 \longrightarrow 3S + 2H_2O$

(c) $CH_2{=}CH_2 + H_2 \longrightarrow CH_3CH_3$

REDOX

There are two ways of finding out whether a substance has been oxidised or reduced during a chemical reaction:

1. electron transfer

2. changes in oxidation state

HALF EQUATIONS

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

This reaction can be divided into two separate equations, one showing oxidation and the other reduction. These are known as half equations.

1. $Na \longrightarrow Na^+ + e^-$ e- is lost

Each sodium atom loses an electron from its outer shell. Sodium atoms have been oxidised.

2. $Cl_2 + 2e^- \longrightarrow 2Cl^-$ e- is gained

Each chlorine atom gains an electron to complete its outer shell. The chlorine atoms have been reduced.

SKILL CHECK

Write two half-equations for the following reactions. For each half-equation state whether oxidation or reduction is occurring:

(a) $Cl_2 + 2I^- \longrightarrow I_2 + 2Cl^-$

(b) $2Mg + O_2 \longrightarrow 2MgO$

(c) $4Fe + 3O_2 \longrightarrow 2Fe_2O_3$

BALANCING HALF EQUATIONS

A balanced ionic equation can be constructed by balancing the number of electrons lost and gained and then adding the two half equations together.

The electrons lost and gained in a redox reaction must be equal.

SKILL CHECK

Zinc metal reacts with IO_3^- ions in acidic solution. Construct a balanced ionic equation for this reaction, using the two half-equations below:

$$2IO_3^- + 12H^+ + 10e^- \longrightarrow I_2 + 6H_2O$$

$$Zn \longrightarrow Zn^{2+} + 2e^-$$

OXIDATION NUMBERS

Oxidation number is the number of electrons that an individual atom within an ion or compound has either lost control of (+ve) or gained control of (-ve).

In simple ionic compounds, the non-metal is assigned a negative oxidation number and the metal is assigned a positive value.

In simple covalent compounds and complex ions, the most electronegative element is assigned a negative value and the least electronegative element is assigned a positive value.

The oxidation number of an uncombined element is 0. e.g. Na, Mg, Cu, O_2, N_2.

OXIDATION NUMBERS

Certain elements (at least in our course) can usually be regarded as having a fixed value in all of their compounds. Examples of these follow.

- Group 1 elements are always +1.
- Group 2 elements are always +2.
- Group 3 elements are always +3.
- Fluorine is always -1.
- Oxygen is nearly always -2 unless it is combined with fluorine or when it is in the form of a peroxide or superoxide.

The total oxidation numbers of all atoms in a neutral compound always add up to 0.

OXIDATION NUMBERS

The oxidation number of an atom in the form of a simple ion is equal to the charge on the ion. e.g.

- The oxidation number of potassium in the ion K^+ = +1
- The oxidation number of aluminium in the ion Al^{3+} = +3
- The oxidation number of nitrogen in the ion N^{3-} = -3
- The oxidation number of oxygen in the ion O^{2-} = -2

The total oxidation numbers of all the atoms in a complex ion always add up to the charge on the ion.

SKILL CHECK

State the oxidation number of the bold atoms in these compounds or ions:

(a) $\mathbf{P}_2O_5$

(b) $\mathbf{S}O_4^{2-}$

(c) $Al_2\mathbf{Cl}_6$

(d) $\mathbf{Cl}O_2^-$

(e) $Ca\mathbf{C}O_3$

OXIDISING AGENT

Increases the oxidation number of another atom.

An atom of the oxidising agent decreases in oxidation numbers.

The oxidising agent gets reduced - it gains electrons.

SKILL CHECK

$2I^- + MnO_2 + 6H^+ + 2SO_4^{2-} \longrightarrow I_2 + Mn^{2+} + 2HSO_4^- + 2H_2O$

Which substance is the oxidising agent? Explain your answer.

REDUCING AGENT

Decreases the oxidation number of another atom.

An atom of the reducing agent increases in oxidation numbers.

The reducing agent gets oxidized - it loses electrons.

SKILL CHECK

$SO_2 + Br_2 + H_2O \longrightarrow SO_4^{2-} + Br^- + H^+$

Identify the reducing agent in this reaction. Explain your answer.

REDOX WS 1

SECTION A

1 When ammonia is converted into nitric acid on a commercial scale, the following reactions can occur.

In which reaction does the greatest change in oxidation number of the nitrogen occur?

reaction	
A	$4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$
B	$3NO_2 + H_2O \rightarrow 2HNO_3 + NO$
C	$2NO + O_2 \rightarrow 2NO_2$
D	$4NH_3 + 6NO \rightarrow 5N_2 + 6H_2O$

2 Granular urea, CON_2H_4, can be used to remove NO_2 from the flue gases of power stations, converting it into harmless nitrogen.

$$2CON_2H_4 + xNO_2 \rightarrow 2CO_2 + yH_2O + zN_2$$

What are the values of *x*, *y* and *z* in a balanced equation?

	x	*y*	*z*
A	1½	2	1¼
B	2	4	3
C	3	4	3½
D	3	4	3

3 A weedkiller can be prepared by heating a bleach solution.

$$\underset{\text{bleach}}{3NaClO} \xrightarrow{\text{heat}} 2NaCl + \underset{\text{weedkiller}}{NaClO_3}$$

What are the oxidation states of chlorine in these three compounds?

A –1 –1 +5

B +1 –1 +5

C +1 –1 +7

D +2 +1 +7

4 The nickel-cadmium rechargeable battery is based upon the following overall reaction.

$$Cd + 2NiOOH + 4H_2O \rightarrow Cd(OH)_2 + 2Ni(OH)_2.H_2O$$

What is the oxidation number of nickel at the beginning and at the end of the reaction?

	beginning	end
A	+1.5	+2
B	+2	+3
C	+3	+2
D	+3	+4

5 The first stage in the manufacture of nitric acid is the oxidation of ammonia by oxygen.

$$\mathbf{w}NH_3(g) + \mathbf{x}O_2(g) \rightarrow \mathbf{y}NO(g) + \mathbf{z}H_2O(g)$$

Which values for **w**, **x**, **y** and **z** are needed to balance the equation?

	w	**x**	**y**	**z**
A	4	5	4	6
B	4	6	4	5
C	5	6	5	4
D	6	5	6	4

6 In an experiment, 50.0 cm^3 of a 0.10 mol dm^{-3} solution of a metallic salt reacted exactly with 25.0 cm^3 of 0.10 mol dm^{-3} aqueous sodium sulphite.

The half-equation for oxidation of sulphite ion is shown below.

$$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

If the original oxidation number of the metal in the salt was +3, what would be the new oxidation number of the metal?

A +1 **B** +2 **C** +4 **D** +5

7 In black and white photographic film, light converts silver chloride into metallic silver. After the film has been developed, the unreacted silver chloride is removed by reaction with sodium thiosulfate to produce a ‘fixed’ negative.

$$AgCl + 2Na_2S_2O_3 \rightarrow 4Na^+ + Cl^- + [Ag(S_2O_3)_2]^{3-}$$

What is the function of the thiosulfate ion?

A to make the silver ions soluble

B to oxidise the silver ions

C to oxidise the silver metal

D to reduce the silver ions

8 Ammonium nitrate, NH_4NO_3, can decompose explosively when heated.

$$NH_4NO_3 \rightarrow N_2O + 2H_2O$$

What are the changes in the oxidation numbers of the two nitrogen atoms in NH_4NO_3 when this reaction proceeds?

A –2, –4 **B** +2, +6 **C** +4, –6 **D** +4, –4

9 Sulfur dioxide, SO_2, is added to wines to prevent oxidation of ethanol by air. To determine the amount of SO_2, a sample of wine is titrated with iodine, I_2. In this reaction, **one** mole of SO_2 is oxidised by **one** mole of I_2.

What is the change in oxidation number of sulfur in this reaction?

A +2 to +4 **B** +2 to +6 **C** +4 to +5 **D** +4 to +6

10 In aqueous solution, the acid HIO disproportionates according to the following equation where **m**, **n**, **p** and **q** are simple whole numbers in their lowest ratios.

$$\mathbf{m}HIO \rightarrow \mathbf{n}I_2 + \mathbf{p}HIO_3 + \mathbf{q}H_2O$$

This equation can be balanced using oxidation numbers.

What are the values for **n** and **p**?

	n	**p**
A	1	2
B	2	1
C	4	1
D	4	2

11 X, Y and Z represent different halogens. The table shows the results of nine experiments in which aqueous solutions of X_2, Y_2 and Z_2 were separately added to separate aqueous solutions containing X^-, Y^- and Z^- ions.

	X^-(aq)	Y^-(aq)	Z^-(aq)
X_2(aq)	no reaction	no reaction	no reaction
Y_2(aq)	X_2 formed	no reaction	Z_2 formed
Z_2(aq)	X_2 formed	no reaction	no reaction

Which row in the following table contains the ions X^-, Y^- and Z^- in order of their decreasing strength as reducing agents?

	strongest	⟶	weakest
A	X^-	Y^-	Z^-
B	X^-	Z^-	Y^-
C	Y^-	Z^-	X^-
D	Z^-	X^-	Y^-

12 When chlorine and aqueous sodium hydroxide are heated together the following overall reaction occurs.

$$3Cl_2(aq) + 6NaOH(aq) \rightarrow 5NaCl(aq) + NaClO_3(aq) + 3H_2O(l)$$

What are the oxidation numbers for chlorine in each of the following species?

	Cl_2	$NaCl$	$NaClO_3$
A	0	+1	–5
B	+2	–1	+3
C	0	–1	+5
D	–2	+1	–3

13 The following equations the letters **W**, **X**, **Y** and **Z** all represent whole numbers.

When correctly balanced, which equation requires one of letters **W**, **X**, **Y** or **Z** to be 5?

A $\mathbf{W}C_3H_7COOH + \mathbf{X}O_2 \rightarrow \mathbf{Y}CO_2 + \mathbf{Z}H_2O$

B $\mathbf{W}C_4H_8 + \mathbf{X}O_2 \rightarrow \mathbf{Y}CO_2 + \mathbf{Z}H_2O$

C $\mathbf{W}H_3PO_4 + \mathbf{X}NaOH \rightarrow \mathbf{Y}Na_2HPO_4 + \mathbf{Z}H_2O$

D $\mathbf{W}NH_3 + \mathbf{X}O_2 \rightarrow \mathbf{Y}N_2 + \mathbf{Z}H_2O$

14 Sodium iodide reacts with concentrated sulfuric acid. The equation which represents one of the reactions that takes place is shown.

$$8NaI + 9H_2SO_4 \rightarrow 8NaHSO_4 + 4I_2 + H_2S + 4H_2O$$

Which species has been oxidised in this reaction?

A H^+ **B** I^- **C** Na^+ **D** SO_4^{2-}

15 The oxide of titanium, TiO_2, is used as a 'whitener' in toothpaste. It is obtained from the ore iron(II) titanate, $FeTiO_3$.

What is the change, if any, in the oxidation number (oxidation state) of titanium in the reaction $FeTiO_3 \rightarrow TiO_2$?

A It is oxidised from +3 to +4.

B It is reduced from +3 to +2.

C It is reduced from +6 to +4.

D There is no change in the oxidation number.

16 In the treatment of domestic water supplies, chlorine is added to the water to form HC*l*O.

$$Cl_2(aq) + H_2O(l) \rightarrow H^+(aq) + Cl^-(aq) + HClO(aq)$$

The HC*l*O reacts further to give ClO^- ions.

$$HClO(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO^-(aq)$$

Both HC*l*O and ClO^- kill bacteria by oxidation.

What is the overall change in oxidation number of chlorine when forming the ClO^- ion from the aqueous chlorine?

A –1 **B** 0 **C** +1 **D** +2

17 In which reaction does an element undergo the largest change in oxidation state?

A $Cl_2 + 2OH^- \rightarrow OCl^- + Cl^- + H_2O$

B $3Cl_2 + 6OH^- \rightarrow ClO_3^- + 5Cl^- + 3H_2O$

C $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

D $3MnO_4^{2-} + 4H^+ \rightarrow MnO_2 + 2MnO_4^- + 2H_2O$

18 The reaction between KI and concentrated H_2SO_4 is a redox reaction.

$$5H_2SO_4 + 8KI \rightarrow 4K_2SO_4 + 4I_2 + H_2S + 4H_2O$$

What is the change in oxidation state of the element that is reduced?

A 1 **B** 4 **C** 6 **D** 8

19 Solutions containing chlorate(I) ions are used as household bleaches and disinfectants. These solutions decompose on heating as shown.

$$3ClO^- \rightarrow ClO_3^- + 2Cl^-$$

Which oxidation state is shown by chlorine in each of these three ions?

	ClO^-	ClO_3^-	Cl^-
A	+1	+3	–1
B	–1	+3	+1
C	+1	+5	–1
D	–1	+5	+1

20 In the redox reaction shown, how do the oxidation states of vanadium and sulfur change?

$$VO_2^+ + SO_2 \rightarrow V^{3+} + SO_4^{2-}$$

	vanadium		sulfur	
	from	to	from	to
A	+1	+3	0	–2
B	+1	+3	+4	+6
C	+5	+3	0	–2
D	+5	+3	+4	+6

21 In which reaction does a single nitrogen atom have the greatest change in oxidation number?

A $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

B $3NO_2 + H_2O \rightarrow 2HNO_3 + NO$

C $2NO + O_2 \rightarrow 2NO_2$

D $4NH_3 + 6NO \rightarrow 5N_2 + 6H_2O$

22 The following half reactions occur when potassium iodate(V), KIO_3, in hydrochloric acid solution oxidises iodine to ICl_2^-.

$$IO_3^- + 2Cl^- + 6H^+ + 4e^- \rightarrow ICl_2^- + 3H_2O$$

$$I_2 + 4Cl^- \rightarrow 2ICl_2^- + 2e^-$$

What is the ratio of IO_3^- to I_2 in the balanced chemical equation for the overall reaction?

A 1 : 1 **B** 1 : 2 **C** 1 : 4 **D** 2 : 1

23 Solid potassium halides react with concentrated sulfuric acid, according to the following equations.

reaction 1 $2KCl + H_2SO_4 \rightarrow K_2SO_4 + 2HCl$

reaction 2 $2KBr + 2H_2SO_4 \rightarrow K_2SO_4 + SO_2 + Br_2 + 2H_2O$

reaction 3 $8KI + 5H_2SO_4 \rightarrow 4K_2SO_4 + H_2S + 4I_2 + 4H_2O$

What is the largest **change** in the oxidation number of sulfur in each of these reactions?

	reaction 1	reaction 2	reaction 3
A	0	0	4
B	0	2	4
C	0	2	8
D	0	4	8

24 Ammonium nitrate, NH_4NO_3, can decompose explosively when heated.

$$NH_4NO_3 \rightarrow N_2O + 2H_2O$$

What are the changes in the oxidation numbers of the two nitrogen atoms in NH_4NO_3 when this reaction proceeds?

A –2, –4 **B** +2, +6 **C** +4, –6 **D** +4, –4

25 In which reaction is the species in **bold** acting as an oxidising agent?

A $\mathbf{2Ca} + O_2 \rightarrow 2CaO$

B $Cr_2O_7^{2-} + 8H^+ + \mathbf{3SO_3^{2-}} \rightarrow 2Cr^{3+} + 4H_2O + 3SO_4^{2-}$

C $Mg + \mathbf{Fe^{2+}} \rightarrow Mg^{2+} + Fe$

D $\mathbf{SO_2} + 2H_2O + 2Cu^{2+} + 2Cl^- \rightarrow H_2SO_4 + 2H^+ + 2CuCl$

26 In which reaction does hydrogen behave as an oxidising agent?

A $H_2 + Cl_2 \rightarrow 2HCl$

B $C_2H_4 + H_2 \rightarrow C_2H_6$

C $N_2 + 3H_2 \rightarrow 2NH_3$

D $2Na + H_2 \rightarrow 2NaH$

27 Chlorine gas reacts with cold aqueous sodium hydroxide. It can also react with hot aqueous sodium hydroxide.

What are the oxidation numbers of chlorine in the products of these reactions?

	cold aqueous sodium hydroxide	hot aqueous sodium hydroxide
A	–1, +1	–1, +5
B	–1, +1	+1, +6
C	–1, +2	–1, +5
D	–1, +2	+1, +6

28 In oxygen difluoride, OF_2, fluorine has an oxidation number of –1. OF_2 will react with sulfur dioxide according to the following equation.

$$OF_2 + SO_2 \rightarrow SO_3 + F_2$$

What is oxidised and what is reduced in this reaction?

	fluorine	oxygen in OF_2	sulfur
A	oxidised	oxidised	reduced
B	oxidised	reduced	oxidised
C	reduced	oxidised	reduced
D	reduced	reduced	oxidised

29 Chlorate(V) ions, ClO_3^-, are produced in the redox reaction between chlorine and hot aqueous sodium hydroxide. Oxidation numbers can be used to help balance the equation for this reaction.

What will be the values of coefficients **v**, **x** and **y** in the balanced equation?

$$\mathbf{v}Cl_2(g) + \mathbf{w}OH^-(aq) \rightarrow \mathbf{x}Cl^-(aq) + \mathbf{y}ClO_3^-(aq) + \mathbf{z}H_2O(l)$$

	v	**x**	**y**
A	2	3	1
B	3	4	2
C	3	5	1
D	7	12	2

30 In which reaction does an element have the largest change in oxidation number?

A $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

B $3OCl^- \rightarrow ClO_3^- + 2Cl^-$

C $5Fe^{2+} + MnO_4^- + 8H^+ \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O$

D $PbO_2 + Sn^{2+} + 4H^+ \rightarrow Sn^{4+} + Pb^{2+} + 2H_2O$

31 Arsenic chloride, $AsCl_3$, reacts with sodium borohydride, $NaBH_4$.

$$\mathbf{p}AsCl_3 + \mathbf{q}NaBH_4 \rightarrow \mathbf{r}AsH_3 + \mathbf{s}NaCl + \mathbf{t}BCl_3$$

What are the numbers **p**, **q**, **r**, **s** and **t** when this equation is balanced correctly?

	p	**q**	**r**	**s**	**t**
A	2	3	2	3	1
B	3	3	3	3	2
C	4	3	4	3	3
D	4	4	4	4	3

32 Which compound contains two different elements with identical oxidation states?

A $HClO$ **B** $Mg(OH)_2$ **C** Na_2SO_4 **D** NH_4Cl

33 In the treatment of domestic water supplies, chlorine is added to the water to form $HClO$.

$$Cl_2(aq) + H_2O(l) \rightarrow H^+(aq) + Cl^-(aq) + HClO(aq)$$

The $HClO$ reacts further to give ClO^- ions.

$$HClO(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + ClO^-(aq)$$

Both $HClO$ and ClO^- kill bacteria by oxidation.

What is the change in oxidation number of chlorine when forming the ClO^- ion from aqueous chlorine?

A –1 **B** 0 **C** +1 **D** +2

34 The reaction between acidified dichromate(VI) ions, $Cr_2O_7^{2-}$, and aqueous Fe^{2+} ions results in the dichromate(VI) ions being reduced to Cr^{3+} ions.

What is the correct equation for this reaction?

A $Cr_2O_7^{2-} + Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + Fe^{3+} + 7H_2O$

B $Cr_2O_7^{2-} + 2Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 2Fe^{3+} + 7H_2O$

C $Cr_2O_7^{2-} + 3Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 3Fe^{3+} + 7H_2O$

D $Cr_2O_7^{2-} + 6Fe^{2+} + 14H^+ \rightarrow 2Cr^{3+} + 6Fe^{3+} + 7H_2O$

35 Sodium chromate(VI), Na_2CrO_4, is manufactured by heating chromite, $FeCr_2O_4$, with sodium carbonate in an oxidising atmosphere. Chromite contains $Cr_2O_4^{2-}$ ions.

$$2FeCr_2O_4 + 4Na_2CO_3 + 3\tfrac{1}{2}O_2 \rightarrow 4Na_2CrO_4 + Fe_2O_3 + 4CO_2$$

What happens in this reaction?

A Chromium and iron are the only elements oxidised.

B Chromium, iron and carbon are oxidised.

C Only chromium is oxidised.

D Only iron is oxidised.

36 When copper reacts with a 50% solution of nitric acid, nitrogen monoxide is evolved and a blue solution results.

The balanced equation for this reaction is shown.

$$p\,\mathrm{Cu} + q\,\mathrm{HNO_3} \rightarrow r\,\mathrm{Cu(NO_3)_2} + s\,\mathrm{H_2O} + t\,\mathrm{NO}$$

What are the values of the integers *p*, *q*, *r*, *s* and *t*?

	p	*q*	*r*	*s*	*t*
A	1	4	1	2	2
B	2	6	2	3	2
C	2	8	2	4	4
D	3	8	3	4	2

37 When K_2MnO_4 is dissolved in water, the following reaction occurs.

$$a\,\mathrm{MnO_4^{2-}(aq)} + b\,\mathrm{H_2O(l)} \rightarrow c\,\mathrm{MnO_4^{-}(aq)} + d\,\mathrm{MnO_2(s)} + e\,\mathrm{OH^{-}(aq)}$$

What are the values of *a* and *c* in the balanced chemical equation?

	a	*c*
A	2	1
B	3	2
C	4	3
D	5	4

38 In some fireworks there is a reaction between powdered aluminium and powdered barium nitrate. Heat is evolved, an unreactive gas is produced, and all nitrogen atoms are reduced.

What is the equation for this reaction?

A $2\mathrm{A}l + \mathrm{Ba(NO_3)_2} \rightarrow \mathrm{A}l_2\mathrm{O_3} + \mathrm{BaO} + 2\mathrm{NO}$

B $4\mathrm{A}l + 4\mathrm{Ba(NO_3)_2} \rightarrow 2\mathrm{A}l_2\mathrm{O_3} + 4\mathrm{Ba(NO_2)_2} + \mathrm{O_2}$

C $10\mathrm{A}l + 3\mathrm{Ba(NO_3)_2} \rightarrow 5\mathrm{A}l_2\mathrm{O_3} + 3\mathrm{BaO} + 3\mathrm{N_2}$

D $10\mathrm{A}l + 18\mathrm{Ba(NO_3)_2} \rightarrow 10\mathrm{A}l\mathrm{(NO_3)_3} + 18\mathrm{BaO} + 3\mathrm{N_2}$

39 A reaction sequence is shown.

$$SO_2 \xrightarrow{1} SO_3 \xrightarrow{2} H_2SO_4 \xrightarrow{3} H_2S \xrightarrow{4} SO_2$$

In each stage of this sequence the sulfur is oxidised, reduced or neither oxidised nor reduced.

Which row is correct?

	1	2	3	4
A	neither	oxidised	reduced	reduced
B	oxidised	neither	reduced	reduced
C	oxidised	neither	reduced	oxidised
D	oxidised	oxidised	reduced	oxidised

40 Vanadium reacts with dilute sulfuric acid to form $V_2(SO_4)_3$ and hydrogen gas.

What happens to vanadium atoms in this reaction?

A They lose three electrons and are oxidised.

B They lose three electrons and are reduced.

C They lose two electrons and are oxidised.

D They lose two electrons and are reduced.

41 $Na_2S_2O_3$ reacts with HCl as shown.

$$Na_2S_2O_3(aq) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l) + SO_2(g) + S(s)$$

When calculating the oxidation number of sulfur in $Na_2S_2O_3$, the **average** oxidation number of the two sulfur atoms should be found.

What is the oxidation number of sulfur in each of $Na_2S_2O_3$, SO_2, and S?

	$Na_2S_2O_3$	SO_2	S
A	+2	+2	+1
B	+2	+4	0
C	+4	+4	0
D	+5	+4	0

42 Which statement is **always** correct for an oxidation reaction?

A It involves the gain of oxygen by an element.

B For one reactant to be oxidised a different reactant must be reduced.

C The element or ion being oxidised will gain electrons.

D The oxidation number of the element being oxidised will increase.

[W’17 2 Q8]

43 One of the reactions in a lead / acid cell is shown.

$$Pb(s) + PbO_2(s) + 4H^+(aq) + 2SO_4^{2-}(aq) \rightarrow 2PbSO_4(s) + 2H_2O(l)$$

Which statement about this reaction is correct?

A Lead is both oxidised and reduced.

B Lead is neither oxidised nor reduced.

C Lead is oxidised only.

D Lead is reduced only.

[W'17 3 Q6]

44 Which reaction is **not** a redox reaction?

A $Mg + 2HNO_3 \rightarrow Mg(NO_3)_2 + H_2$

B $2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$

C $SO_2 + NO_2 \rightarrow SO_3 + NO$

D $SO_3 + H_2O \rightarrow H_2SO_4$

[M'18 2 Q10]

45 When burned, sulfur forms a gaseous product X which can be oxidised to produce a gas Y.

Gas Y reacts with water to produce a product Z.

Which row correctly shows the oxidation states of sulfur in X, Y and Z?

	X	Y	Z
A	–2	+4	+4
B	–2	+4	+6
C	+4	+6	+4
D	+4	+6	+6

[M'18 2 Q18]

46 Ethanedioate ions, $C_2O_4^{2-}$, react with a suitable reagent to form CO_2. A half-equation for this reaction is shown.

$$C_2O_4^{2-} \rightarrow 2CO_2 + 2e^-$$

Which row is correct?

	oxidation state of carbon in $C_2O_4^{2-}$	type of reaction
A	+3	oxidation
B	+3	reduction
C	+5	oxidation
D	+5	reduction

[S'18 1 Q8]

47 Oxidation numbers should be used to answer this question.

A redox reaction takes place between hydroxylammonium ions, $[NH_3OH]^+$, and acidified iron(III) ions, Fe^{3+}. The products are iron(II) ions, Fe^{2+}, H^+ ions, water and a compound of nitrogen.

The mole ratio of reacting hydroxylammonium ions to reacting iron(III) ions is 1 : 2.

Which nitrogen-containing compound could be formed in the reaction?

A NH_3 **B** N_2O **C** NO **D** NO_2

[S'18 1 Q9]

48 Sulfur reacts with concentrated nitric acid in a redox reaction.

$$S + 4HNO_3 \rightarrow SO_2 + 4NO_2 + 2H_2O$$

What are the changes in oxidation number of sulfur and of nitrogen in this reaction?

	sulfur	nitrogen
A	+2	–3
B	+2	–1
C	+4	–3
D	+4	–1

[S'18 2 Q8]

49 Xenon hexafluoride, XeF_6, reacts with water.

$$XeF_6 + 3H_2O \rightarrow XeO_3 + 6HF$$

Which statement is correct?

A Hydrogen is reduced in this reaction.

B Hydrogen is the only element oxidised in this reaction.

C The only element oxidised in this reaction is xenon.

D This is not a redox reaction.

[S'18 3 Q8]

SECTION B

The responses **A** to **D** should be selected on the basis of

A	B	C	D
1, **2** and **3** are correct	**1** and **2** only are correct	**2** and **3** only are correct	**1** only is correct

1 Chlorine reacts with hot concentrated aqueous sodium hydroxide according to the equation below.

$$3Cl_2(g) + 6NaOH(aq) \rightarrow NaClO_3(aq) + 5NaCl(aq) + 3H_2O(l)$$

Which conclusions can be drawn from this information?

1 The oxidation state of the chlorine in one of the products is +5.

2 The chlorine undergoes disproportionation.

3 The sodium hydroxide acts as a reducing agent.

2 Disproportionation is the term used to describe a reaction in which a reactant is simultaneously both oxidised and reduced.

To which incomplete equations does the term disproportionation apply?

1 $Cl_2(g) + 2OH^-(aq) \rightarrow H_2O(l) + Cl^-(aq) + \ldots\ldots$

2 $3Cl_2(g) + 6OH^-(aq) \rightarrow 3H_2O(l) + ClO_3^-(aq) + \ldots\ldots$

3 $2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + \ldots\ldots$

3 Zirconium, Zr, proton number 40, is a metal which is used in corrosion-resistant alloys.

Zirconium metal is extracted from the oxide ZrO_2 by the following sequence of reactions.

reaction 1 $ZrO_2 + 2Cl_2 + 2C \rightarrow ZrCl_4 + 2CO$

reaction 2 $ZrCl_4 + 2Mg \rightarrow Zr + 2MgCl_2$

Which statements about this extraction process are correct?

1 Carbon in reaction 1 behaves as a reducing agent.

2 Magnesium in reaction 2 behaves as a reducing agent.

3 Chlorine in reaction 1 behaves as a reducing agent.

4 A space shuttle's upward thrust came from the following reaction between aluminium and ammonium perchlorate.

$$10Al + 6NH_4ClO_4 \rightarrow 4Al_2O_3 + 2AlCl_3 + 12H_2O + 3N_2$$

Which statements about this reaction are correct?

1 Aluminium is oxidised.

2 Chlorine is reduced.

3 Nitrogen is oxidised.

5 Many crude oils contain H_2S. During refining, by the Claus process, the H_2S is converted into solid sulfur, which is then removed.

reaction I $2H_2S(g) + 3O_2(g) \rightarrow 2H_2O(l) + 2SO_2(g)$

reaction II $2H_2S(g) + SO_2(g) \rightarrow 2H_2O(l) + 3S(s)$

Which statements about the Claus process are correct?

1 H_2S is oxidised in reaction I.

2 SO_2 oxidises H_2S in reaction II.

3 Hydrogen is oxidised in reaction II.

6 In a solution that contains both Br_2 and Cl_2, a process takes place that produces BrO_3^- ions.

The process is represented by the following equations.

equation 1 $Br_2 + H_2O \rightarrow HBr + HBrO$

equation 2 $3HBrO + Cl_2 \rightarrow 2Cl^- + BrO_3^- + Br_2 + 3H^+$

Which statements about these reactions are correct?

1 Chlorine is reduced in equation 2.

2 Bromine is oxidised in both equation 1 and equation 2.

3 Bromine is reduced in both equation 1 and equation 2.

7 Which of these reactions are redox reactions?

1 $6NO(g) + 4NH_3(g) \rightarrow 5N_2(g) + 6H_2O(g)$

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

3 $SO_3(g) + H_2O(g) \rightarrow H_2SO_4(g)$

8 The salt $NaClO_3$ is used as a non-selective weedkiller.

On careful heating, this reaction occurs: $4NaClO_3 \rightarrow NaCl + 3NaClO_4$.

On strong heating this reaction occurs: $NaClO_4 \rightarrow NaCl + 2O_2$.

The overall reaction is $2NaClO_3 \rightarrow 2NaCl + 3O_2$.

What do these equations show?

1 $NaClO_3$ can behave as an oxidising agent.

2 $NaClO_3$ can behave as a reducing agent.

3 The oxidation numbers of chlorine in the three compounds shown are +6, +8 and –1.

9 Bromine reacts with water.

$$Br_2 + H_2O \rightleftharpoons HOBr + HBr$$

Which oxidation states of bromine are present in the equilibrium mixture?

1 +3

2 0

3 –1

10 When $KClO_3$ is heated, the following reaction occurs.

$$4KClO_3 \rightarrow 3KClO_4 + KCl$$

Which statements are correct?

1 The oxidation state of Cl in $KClO_3$ is +5.

2 The oxidation state of some Cl atoms decreases by 6.

3 The reaction involves disproportionation.

11 Hydrogen sulfide can be oxidised to form sulfur dioxide.

$$2H_2S + 3O_2 \rightarrow 2SO_2 + 2H_2O$$

Which statements are correct?

1 The oxidation number of sulfur increases by 6.

2 The oxidation number of oxygen increases by 2.

3 The oxidation number of hydrogen decreases.

[W'17 2 Q33]

12 One way of recovering tin from old printed circuit boards is to dissolve it in a mixture of concentrated hydrochloric acid and concentrated nitric acid. The tin dissolves because it reacts with the mixture of these concentrated acids.

$$Sn + 4HCl + 2HNO_3 \rightarrow SnCl_4 + NO_2 + NO + 3H_2O$$

Which statements about this reaction are correct?

1 Nitrogen is present in three different oxidation states in the reactants and products.

2 The oxidation state of tin increases from 0 to +4.

3 The oxidation state of chlorine remains the same.

[W'17 3 Q33]

NOTES

NOTES

REDOX WS 2

1 Biological fixation of nitrogen from the air can occur in three stages in soil, producing nitrate ions, NO_3^-. Protein is synthesised by the plant from nitrate ions.

Stage 1

At the root nodules of plants, nitrogen is converted into ammonium ions.

$$N_2(g) + 8H^+(aq) + 6e^- \longrightarrow 2NH_4^+(aq)$$

Stage 2

The action of aerobic bacteria oxidises ammonium ions to nitrite ions.

$$2NH_4^+(aq) + 3O_2(g) \longrightarrow 2NO_2^-(aq) + 4H^+(aq) + 2H_2O(l)$$

Stage 3

The action of other bacteria oxidises nitrite to nitrate ions.

$$2NO_2^-(aq) + O_2(g) \longrightarrow 2NO_3^-(aq)$$

(a) State the oxidation states of nitrogen in the four nitrogen containing species in this sequence.

N_2 .... NO_2^- ....

NH_4^+ .... NO_3^- .... [2]

2 Antimony, Sb, has been known for about 6000 years. It is present in many ancient forms of bronze, but now its main use is to strengthen lead alloys.

Antimony is produced in a two-stage process from stibnite, a sulphide ore, Sb_2S_3.

The ore is first roasted in oxygen to form the oxide.

$$2Sb_2S_3 + \text{..............}\ O_2 \rightarrow Sb_4O_6 + \text{..............}SO_2$$

(i) Balance the above equation.

The oxide is then reduced with carbon.

$$Sb_4O_6 + 3C \rightarrow 4Sb + 3CO_2$$

(ii) What is the oxidation number of antimony in Sb_4O_6?

....

3 Compounds of phosphorus have many uses in everyday life, e.g. fertilisers, matches and in water softeners.

(a) State the full electronic configuration of phosphorus.

....[1]

(b) Phosphoric acid, H_3PO_4, is used in the manufacture of phosphate fertilisers.

Deduce the oxidation number of phosphorus in H_3PO_4.

....[1]

4 Ammonia does not burn in air but will burn in pure oxygen.

(i) Balance the equation for this reaction:

$$\ldots\ldots\ NH_3(g) + \ldots\ldots\ O_2(g) \longrightarrow \ldots\ldots\ N_2(g) + \ldots\ldots\ H_2O(g)$$

(ii) Use oxidation numbers to explain why this is a redox reaction.

....

....

....[3]

5 The metals magnesium and aluminium form hydrides with formulae MgH_2 and AlH_3. The non-metals phosphorus and sulfur form hydrides with formulae PH_3 and H_2S.
By considering their positions in the Periodic Table, suggest oxidation numbers for these four elements in their hydrides.

compound	MgH_2	AlH_3	PH_3	H_2S
oxidation number of element in the hydride				

6 **(d)** Mohr's salt, $(NH_4)_2Fe(SO_4)_2•xH_2O$, is the hydrated form of ammonium iron(II) sulfate.

x represents the number of moles of water in 1 mole of the salt.

A student wanted to determine the value of x. 0.784 g of the hydrated salt was dissolved in water and this solution was acidified.

All of the solution was titrated with 0.0200 mol dm^{-3} potassium manganate(VII). 20.0 cm^3 of this potassium manganate(VII) solution was required for complete reaction with the Fe^{2+} ions.

(i) Use changes in oxidation numbers to balance the equation for the reaction taking place.

$MnO_4^-(aq)$ + ...$Fe^{2+}(aq)$ + ...$H^+(aq)$ → .....$Mn^{2+}(aq)$ + ...$Fe^{3+}(aq)$ + ...$H_2O(l)$

[1]

(ii) State the role of the Fe^{2+} ions in this reaction.

Explain your answer.

....

.... [2]

(iii) Calculate the amount, in moles, of manganate(VII) ions that reacted.

amount = ... mol [1]

(iv) Calculate the amount, in moles, of Fe^{2+} ions in the sample of the salt.

amount = ... mol [1]

(v) Calculate the relative formula mass of $(NH_4)_2Fe(SO_4)_2•xH_2O$.

relative formula mass = ... [1]

(vi) Calculate the value of x.

x = ... [1]

[S’18 2 Q2]

NOTES

Group 2

The physical and chemical properties of the elements of Group 2 (the alkaline Earth metals) are introduced in this topic.

10.1 Similarities and trends in the properties of the Group 2 metals, magnesium to barium, and their compounds ☐

10.2 Some uses of Group 2 compounds ☐

GROUP 2

10 Group 2

The physical and chemical properties of the elements of Group 2 (the alkaline Earth metals) are introduced in this topic.

	Learning outcomes Candidates should be able to:
10.1 Similarities and trends in the properties of the Group 2 metals, magnesium to barium, and their compounds	a) describe the reactions of the elements with oxygen, water and dilute acids b) describe the behaviour of the oxides, hydroxides and carbonates with water and dilute acids c) describe the thermal decomposition of the nitrates and carbonates d) interpret, and make predictions from, the trends in physical and chemical properties of the elements and their compounds e) state the variation in the solubilities of the hydroxides and sulfates **f) interpret and explain qualitatively the trend in the thermal stability of the nitrates and carbonates in terms of the charge density of the cation and the polarisability of the large anion** **g) interpret and explain qualitatively the variation in solubility of the hydroxides and sulfates in terms of relative magnitudes of the enthalpy change of hydration and the corresponding lattice energy**
10.2 Some uses of Group 2 compounds	a) describe and explain the use of calcium hydroxide and calcium carbonate (powdered limestone) in agriculture

INTRODUCTION

The Group II metals (also known as earth metals) are all elements with similar properties.

Earth Metals comprise of Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra).

They are generally shiny, silvery-white metals with relatively low densities and melting points.

Beryllium (Be)	$1s^2\ 2s^2$
Magnesium (Mg)	$1s^2\ 2s^2\ 2p^6\ 3s^2$
Calcium (Ca)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2$
Strontium (Sr)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 5s^2$
Barium (Ba)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ 4d^{10}\ 5s^2\ 5p^6\ 6s^2$

PHYSICAL PROPERTIES

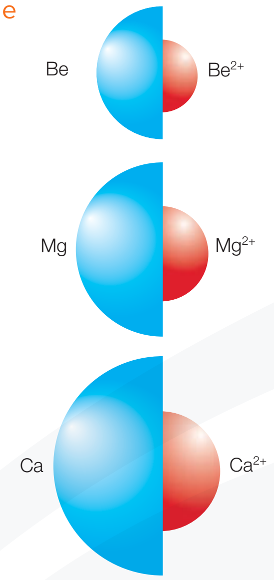
	Mg	Ca	Sr	Ba
Melting point/°C	649	839	769	729
Boiling point/°C	1090	1484	1384	1637
First ionisation energy/kJ mol^{-1}	736	590	548	502
Second ionisation energy/kJ mol^{-1}	1450	1150	1060	966
Ionic radius of M^{2+} ion/nm	0.065	0.099	0.113	0.135

Group 2 metals are reducing agents. They readily give up their two s electrons to form M^{2+} ions (where M represents Mg, Ca, Sr or Ba).

The elements get more reactive as we go down the group as it takes less energy to remove the pair of outer electrons going down Group 2.

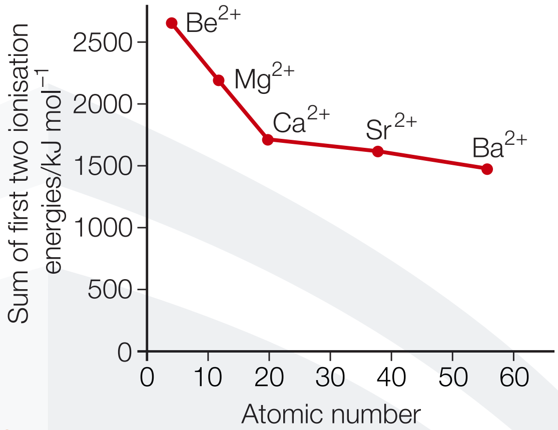
PHYSICAL PROPERTIES

Relative sizes of the atoms and ions of Group 2 elements.



Graph to show the trend in the sum of the first two ionisation energies of Group 2 metals:

$$X(g) \longrightarrow X^{2+}(g) + 2e^-$$



REACTIONS WITH WATER AND STEAM

The metals Mg to Ba in Group 2 react with water. The reactions are not as vigorous as the reactions of the Group 1 metals, but, as in Group 1, the rate of reaction increases down the group.

Magnesium reacts very slowly with cold water but much more rapidly on heating in steam. It does not give the hydroxide, as magnesium oxide is almost insoluble in water.

$$Mg_{(s)} + H_2O_{(g)} \longrightarrow MgO_{(s)} + H_{2\,(g)}$$

Calcium reacts with cold water to produce hydrogen and calcium hydroxide.

$$Ca_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2\,(s)/(aq)} + H_{2\,(g)}$$

Barium reacts even faster with cold water, but its hydroxide is more soluble.

REACTIONS WITH WATER AND STEAM

General Reaction: $X + 2H_2O \longrightarrow X(OH)_2 + H_2$

Element	Water	Steam
Magnesium	Slow reaction. Bubble of hydrogen form. Produces a hydroxide and hydrogen.	Burns in steam rapidly. Produces white an oxide and hydrogen.
Calcium, Strontium & Barium	Rapid reactions. Vigor increases down the group. Alkaline solution obtained. Produces a hydroxide and hydrogen.	Explosive reaction. Produces white oxide and hydrogen.

SKILL CHECK

Write balanced equations, including state symbols, for the reaction of:

(a) strontium with water

(b) barium with water

REACTIONS WITH OXYGEN

Apart from beryllium, the Group 2 metals burn brightly on heating in oxygen to form white ionic oxides, $M^{2+}O^{2-}$.

Magnesium burns very brightly in air, with an intense white flame, to form the white solid magnesium oxide, MgO. For this reason, magnesium powder is an ingredient of fireworks and flares.

$$Mg_{(s)} + \frac{1}{2}O_{2\,(g)} \longrightarrow MgO_{(s)}$$

Calcium also burns brightly in air, but with a red flame, to form the white solid calcium oxide, CaO.

Strontium burns with a crimson flame and barium with a green flame.

REACTIONS WITH ACIDS

General Reaction: metal + acid ⟶ hydrogen + salt

All Group 2 Metals react with acids to liberate Hydrogen gas. The silver metal is seen dissolving in the metal.

Reactivity increases down the group.

A soluble salt is produced, except in sulfuric acid.

This is due to sulfates being increasingly insoluble down the group.

GROUP 2 OXIDES

Earth Metals form basic oxides. They have high melting points and are used industrially.

They are normally prepared by heating the hydroxide or carbonate to release water or carbon dioxide gas.

$$Mg(OH)_2 \longrightarrow MgO + H_2O$$

$$MgCO_3 \longrightarrow MgO + CO_2$$

GENERAL PROPERTIES OF COMPOUNDS

Group II Hydroxides and Oxides are increasingly soluble down the group, making alkaline solutions.

Group II Sulfates are decreasingly soluble down the group.

Group II Carbonates are all insoluble in water, and react with acids to liberate carbon dioxide.

$$MgCO_3 + 2HCl \longrightarrow MgCl_2 + CO_2 + H_2O$$

SKILL CHECK

Write a balanced equation, including state symbols, for the reaction of barium carbonate with dilute nitric acid.

THERMAL DECOMPOSITION OF CARBONATES

All Carbonates undergo thermal decomposition to give metal oxide and carbon dioxide gas.

$$MgCO_3 \longrightarrow MgO + CO_2$$

Down the group, the carbonates need to be heated strongly to decompose.

Thermal stability of group 2 carbonates increases down the group.

THERMAL DECOMPOSITION OF NITRATES

All Nitrates undergo thermal decomposition to give metal oxide, nitrogen dioxide (brown gas) and oxygen gas.

$$Mg(NO_3)_2 \longrightarrow MgO + 2\,NO_2 + \tfrac{1}{2}O_2$$

Down the group, nitrates also become increasingly stable to heat.

Thermal stability of group 2 nitrates increases down the group.

SKILL CHECK

Which of the three compounds listed will decompose at the lowest temperature?

(a) calcium carbonate, strontium carbonate, barium carbonate

Q. Write a balanced chemical equation, including state symbols, for the thermal decomposition of:

(a) strontium carbonate

SKILL CHECK

Which of the three compounds listed will decompose at the lowest temperature?

(a) barium nitrate, calcium nitrate, magnesium nitrate

Write a balanced chemical equation, including state symbols, for the thermal decomposition of:

(a) barium nitrate

INDUSTRIAL USES

Magnesium Oxide is used to line blast furnaces. It is highly heat resistant and can be used in furnaces where acid isn't present.

Lime (powdered Calcium Carbonate), Quicklime (Calcium Oxide) and Slaked Lime (Calcium Hydroxide) are used to raise the pH of acidic soils by neutralising the acids.

GROUP 2 WS 1

SECTION A

1 Which equation represents the reaction that occurs when calcium nitrate is heated strongly?

A $Ca(NO_3)_2 \rightarrow Ca(NO_2)_2 + O_2$

B $Ca(NO_3)_2 \rightarrow CaO + N_2O + 2O_2$

C $Ca(NO_3)_2 \rightarrow CaO_2 + 2NO_2$

D $2Ca(NO_3)_2 \rightarrow 2CaO + 4NO_2 + O_2$

2 River water in a chalky agricultural area may contain Ca^{2+}, Mg^{2+}, CO_3^{2-}, HCO_3^-, Cl^-, and NO_3^- ions. In a waterworks, such water is treated by adding a calculated quantity of calcium hydroxide.

What will be precipitated following the addition of calcium hydroxide?

A $CaCl_2$

B $CaCO_3$

C $MgCO_3$

D $Mg(NO_3)_2$

3 What are the products of the thermal decomposition of magnesium nitrate?

A magnesium nitride and oxygen

B magnesium oxide and nitrogen

C magnesium oxide, nitrogen and oxygen

D magnesium oxide, nitrogen dioxide and oxygen

4 The metals of Group II react readily with oxygen to form compounds of general formula *M*O.

When each of these oxides is added to water, which forms the most alkaline solution?

A MgO **B** CaO **C** SrO **D** BaO

5 One mole of each of the following compounds is strongly heated and any gas produced is collected at room temperature and pressure.

From which compound is $24\,dm^3$ of gas likely to be collected?
[One mole of any gas occupies $24\,dm^3$ at room temperature and pressure.]

A $MgCl_2$ **B** $MgCO_3$ **C** $Mg(NO_3)_2$ **D** $Mg(OH)_2$

6 Steam is passed over heated magnesium to give compound **X** and hydrogen.

What is **not** a property of compound **X**?

A It has a high melting point.

B It is a basic oxide.

C It is a white solid.

D It is very soluble in water.

7 A 5.00 g sample of an anhydrous Group II metal nitrate loses 3.29 g in mass on strong heating.

Which metal is present?

A magnesium

B calcium

C strontium

D barium

8 Slaked lime, $Ca(OH)_2$, may be made from limestone, $CaCO_3$.

On heating in a lime kiln at 1000 °C, limestone decomposes as follows.

reaction 1 $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

Water is then reacted with calcium oxide, CaO, as follows.

reaction 2 $CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$

What are the enthalpy changes of these reactions?

	reaction 1	reaction 2
A	endothermic	endothermic
B	endothermic	exothermic
C	exothermic	endothermic
D	exothermic	exothermic

9 Which mass of solid residue can be obtained from the thermal decomposition of 4.10 g of anhydrous calcium nitrate?

A 0.70 g **B** 1.00 g **C** 1.40 g **D** 2.25 g

10 Equimolar quantities of magnesium carbonate and strontium carbonate are separately heated to bring about complete thermal decomposition. The minimum temperature for this to occur is called T_d.

The cold residues are separately added to equal volumes of water and the change in pH is measured. The change in pH is called ΔpH.

Which metal has the higher value of T_d, and the greater value of ΔpH?

	T_d	ΔpH
A	Mg	Mg
B	Mg	Sr
C	Sr	Mg
D	Sr	Sr

11 Steam is passed over heated magnesium to give compound X and hydrogen.

What is **not** a property of compound X?

A It has an M_r of 40.3.

B It is basic.

C It is a white solid.

D It is very soluble in water.

12 When magnesium nitrate, $Mg(NO_3)_2.7H_2O$, is heated, which three gases are given off?

A dinitrogen oxide, oxygen, water vapour

B hydrogen, nitrogen, oxygen

C hydrogen, nitrogen dioxide, oxygen

D nitrogen dioxide, oxygen, water vapour

13 Which property of beryllium and its compounds is typical of the elements below it in Group II?

A Be does not react with hot water.

B $BeCl_2$ is covalent.

C $Be(NO_3)_2$ produces BeO on thermal decomposition.

D BeO dissolves in alkalis.

14 2.920 g of a Group II metal, **X**, reacts with an excess of chlorine to form 5.287 g of a compound with formula **X**Cl_2.

What is metal **X**?

A barium

B calcium

C magnesium

D strontium

15 River water in a chalky agricultural area may contain Ca^{2+}, Mg^{2+}, CO_3^{2-}, HCO_3^-, Cl^- and NO_3^- ions. In a waterworks, such water is treated by adding a calculated quantity of calcium hydroxide.

What will be precipitated following the addition of calcium hydroxide?

A $CaCl_2$

B $CaCO_3$

C $Ca(NO_3)_2$

D $Mg(NO_3)_2$

16 Due to their similar ionic radii, the reactions of lithium and magnesium and their corresponding compounds are very similar.

Which statement concerning the reactions of lithium and its compounds is correct?

A Lithium carbonate decomposes on heating at a relatively low temperature, forming lithium oxide and carbon dioxide.

B Lithium nitrate decomposes on heating, forming lithium nitrite and oxygen.

C Lithium burns only slowly in oxygen.

D Lithium reacts violently with cold water, liberating hydrogen.

17 Which property of Group II elements (beryllium to barium) decreases with increasing atomic number?

A reactivity with water

B second ionisation energy

C solubility of hydroxides

D stability of the carbonates

18 Total elimination of the pollutant sulphur dioxide, SO_2, is difficult, both for economic and technical reasons. Its emission can be reduced in furnace chimneys using desulphurisation plants, where the gases are scrubbed (washed) with calcium hydroxide to remove the SO_2.

What is the main product formed initially?

A CaO **B** $Ca(OH)_2$ **C** $CaSO_3$ **D** $CaSO_4$

19 Which element is expected to show the greatest tendency to form some covalent compounds?

A aluminium

B calcium

C magnesium

D sodium

20 What mass of solid residue can be obtained from the thermal decomposition of 4.10 g of anhydrous calcium nitrate?

A 0.70 g **B** 1.00 g **C** 1.40 g **D** 2.25 g

21 When a mineral was heated in a Bunsen flame to constant mass, a colourless gas that turned lime water milky was evolved. The remaining solid was cooled and then added to aqueous hydrochloric acid. Vigorous effervescence was seen.

What was the mineral?

A aragonite, $CaCO_3$

B artinite, $MgCO_3.Mg(OH)_2.3H_2O$

C barytocalcite, $BaCO_3.CaCO_3$

D dolomite, $CaCO_3.MgCO_3$

22 Limestone, $CaCO_3$, has been used as a building material for thousands of years, and was used on the Pyramids in Egypt. In the past hundred years many limestone buildings have begun to suffer damage.

What is the cause of this damage?

A hydrocarbon emissions from motor vehicles

B increased temperature due to global warming

C increased ultraviolet radiation as the ozone layer is destroyed

D sulphur dioxide from fossil fuels forming 'acid rain'

23 The chemical properties of an element at the top of a group in the Periodic Table are often different from those of the rest of the elements in the group.

Of the following properties of beryllium and its compounds, which property **is** typical of the elements below it in Group II?

A Be does not react with hot water.

B $BeCl_2$ is covalent.

C $Be(NO_3)_2$ produces BeO on thermal decomposition.

D BeO dissolves in alkalis.

24 The combustion of fossil fuels is a major source of increasing atmospheric carbon dioxide, with a consequential rise in global warming. Another significant contribution to carbon dioxide levels comes from the thermal decomposition of limestone, in the manufacture of cement and of lime for agricultural purposes.

Cement works roast 1000 million tonnes of limestone per year and a further 200 million tonnes is roasted in kilns to make lime.

What is the total annual mass output of carbon dioxide (in million tonnes) from these two processes?

A 440 **B** 527 **C** 660 **D** 880

25 What volume of oxygen, measured under room conditions, can be obtained from the thermal decomposition of 8.2 g of calcium nitrate (M_r = 164)?

A 150 cm^3 **B** 300 cm^3 **C** 600 cm^3 **D** 1200 cm^3

26 Lime, CaO, is used to reduce the acidity of soil, and ammonium sulphate is a nitrogenous fertiliser.

Why cannot they be used in a mixed form?

A The dry mixture is explosive.

B $CaSO_4$, formed on mixing, causes hard water.

C When dampened, ammonia is given off.

D Sulphuric acid will form.

27 Which is true for calcium or its compounds compared with the corresponding statements for magnesium?

A Calcium has a smaller atomic radius.

B Calcium oxide reacts less vigorously with water.

C Calcium reacts more vigorously with water.

D The sum of the first two ionisation energies of calcium is greater.

28 Which element is likely to have an electronegativity similar to that of aluminium?

A barium

B beryllium

C magnesium

D strontium

29 Due to their similar ionic radii, the reactions of lithium and magnesium and their corresponding compounds are very similar.

Which statement concerning the reactions of lithium and its compounds is correct?

A Lithium carbonate decomposes on heating at a relatively low temperature, forming lithium oxide and carbon dioxide.

B Lithium nitrate decomposes on heating, forming lithium nitrite and oxygen.

C Lithium only burns slowly in oxygen.

D Lithium reacts violently with cold water, liberating hydrogen.

30 Lime mortar is made from quicklime, water and sand. Over a period of time, lime mortar changes into a much harder form. Both fresh and old lime mortar react with aqueous hydrochloric acid but only the old lime mortar effervesces during the reaction.

Which equation describes the change from fresh to old lime mortar?

A $CaO + CO_2 \rightarrow CaCO_3$

B $CaO + H_2O \rightarrow Ca(OH)_2$

C $Ca(OH)_2 \rightarrow CaO + H_2O$

D $Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$

31 The reaction between aluminium powder and anhydrous barium nitrate is used as the propellant in some fireworks. The metal oxides and nitrogen are the only products.

Which volume of nitrogen, measured under room conditions, is produced when 0.783 g of anhydrous barium nitrate reacts with an excess of aluminium?

A 46.8 cm^3 **B** 72.0 cm^3 **C** 93.6 cm^3 **D** 144 cm^3

32 The oxides BaO, CaO, MgO and SrO all produce alkaline solutions when added to water.

Which oxide produces the saturated solution with the highest pH?

A BaO(aq) **B** CaO(aq) **C** MgO(aq) **D** SrO(aq)

33 What can be seen when a piece of magnesium ribbon is placed in cold water?

A A vigorous effervescence occurs.

B Bubbles of gas form slowly on the magnesium.

C The magnesium floats on the surface of the water and reacts quickly.

D The magnesium glows and a white solid is produced.

34 What are the trends in the stated properties as Group II is descended from magnesium to barium?

	decomposition temperature of the carbonate	first ionisation energy
A	decreases	decreases
B	decreases	increases
C	increases	decreases
D	increases	increases

35 The nitrates of beryllium, calcium, magnesium, and strontium all decompose in the same way when heated. When 2.00 g of one of these anhydrous nitrates is decomposed, 1.32 g of gas is produced.

What is the nitrate?

A beryllium nitrate

B calcium nitrate

C magnesium nitrate

D strontium nitrate

36 When 3.00 g of an anhydrous nitrate of a Group II metal is decomposed, 1.53 g of gas is produced.

What is the nitrate compound?

A beryllium nitrate

B calcium nitrate

C magnesium nitrate

D strontium nitrate

37 Rat poison needs to be insoluble in rain water but soluble at the low pH of stomach contents.

What is a suitable barium compound to use for rat poison?

A barium carbonate

B barium chloride

C barium hydroxide

D barium sulfate

38 In which row of the table are all statements comparing the compounds of magnesium and barium correct?

	solubility of hydroxides		solubility of sulfates	
	solubility of magnesium hydroxide	solubility of barium hydroxide	solubility of magnesium sulfate	solubility of barium sulfate
A	higher	lower	higher	lower
B	higher	lower	lower	higher
C	lower	higher	higher	lower
D	lower	higher	lower	higher

39 Magnesium nitrate, $Mg(NO_3)_2$, will decompose when heated to give a white solid and a mixture of gases. One of the gases released is an oxide of nitrogen, X.

7.4 g of anhydrous magnesium nitrate is heated until no further reaction takes place.

What mass of X is produced?

A 1.5 g **B** 2.3 g **C** 3.0 g **D** 4.6 g

40 Magnesium nitrate, $Mg(NO_3)_2$, will decompose when heated to give a white solid and a mixture of gases. One of the gases released is oxygen.

29.7 g of anhydrous magnesium nitrate is heated until no further reaction takes place.

What mass of oxygen is produced?

A 3.2 g **B** 6.4 g **C** 12.8 g **D** 19.2 g

41 In which row of the table are all statements comparing the compounds of calcium and barium correct?

	solubility of calcium hydroxide	solubility of barium hydroxide	thermal stability of calcium carbonate	thermal stability of barium carbonate
A	higher	lower	higher	lower
B	higher	lower	lower	higher
C	lower	higher	higher	lower
D	lower	higher	lower	higher

42 Anhydrous magnesium nitrate, $Mg(NO_3)_2$, will decompose when heated, giving a white solid and a mixture of two gases **X** and **Y**.

Y is oxygen.

What is the ratio $\frac{\text{mass of } \mathbf{X} \text{ released}}{\text{mass of } \mathbf{Y} \text{ released}}$?

A $\frac{1}{0.174}$ **B** $\frac{1}{0.267}$ **C** $\frac{1}{0.348}$ **D** $\frac{1}{3.43}$

43 In which row of the table are all statements comparing magnesium and barium correct?

	fourth ionisation energy of magnesium	fourth ionisation energy of barium	reaction of magnesium with cold water	reaction of barium with cold water
A	higher	lower	faster	slower
B	higher	lower	slower	faster
C	lower	higher	faster	slower
D	lower	higher	slower	faster

44 A metal, **X**, reacts with water to produce a colourless solution which gives a white precipitate when mixed with aqueous sulfuric acid.

What is metal **X**?

A barium

B magnesium

C potassium

D sodium

45 Which property **increases** in value going down Group II?

A electronegativity

B ionic radius

C maximum oxidation number

D second ionisation energy

46 Which calcium compound contains 54.1 % by mass of calcium?

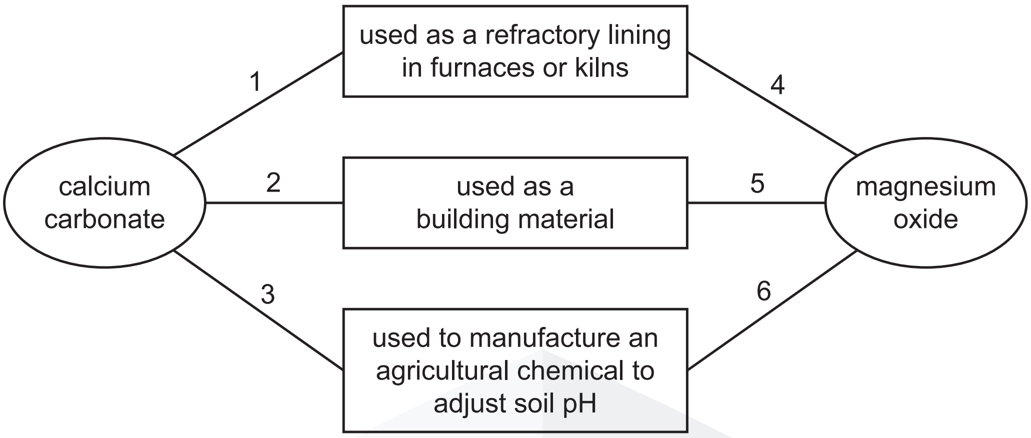
A calcium hydroxide

B calcium nitrate

C calcium oxide

D calcium sulfate

47 The diagram shows some applications of compounds of Group II elements.



Which numbered links are correct?

	calcium carbonate	magnesium oxide
A	1, 2 and 3	4 and 5 only
B	1, 2 and 3	5 and 6 only
C	2 and 3 only	4 only
D	2 and 3 only	6 only

48 River water in a chalky agricultural area may contain Ca^{2+}, Mg^{2+}, CO_3^{2-}, HCO_3^-, Cl^- and NO_3^- ions. In a water treatment plant, such water is treated by adding a calculated quantity of calcium hydroxide.

What will be precipitated from the river water following the addition of calcium hydroxide?

A $CaCl_2$ **B** $CaCO_3$ **C** $Ca(NO_3)_2$ **D** $Mg(NO_3)_2$

49 When equal volumes of saturated solutions of barium hydroxide and calcium hydroxide are mixed, a white precipitate, **Y**, forms. The mixture is filtered and carbon dioxide is bubbled through the filtrate, producing a second white precipitate, **Z**.

What are **Y** and **Z**?

	Y	**Z**
A	$Ba(OH)_2$	$Ca(OH)_2$
B	$Ba(OH)_2$	$CaCO_3$
C	$Ca(OH)_2$	$BaCO_3$
D	$Ca(OH)_2$	$Ba(OH)_2$

50 When calcium is burnt in oxygen, what colour is the flame?

A green

B red

C white

D yellow

51 Group II nitrates undergo thermal decomposition according to the following equation.

$$X(NO_3)_2 \rightarrow XO + 2NO_2 + \tfrac{1}{2}O_2$$

Which Group II nitrate requires the highest temperature to bring about its thermal decomposition?

A barium nitrate

B calcium nitrate

C magnesium nitrate

D strontium nitrate

52 When barium is burnt in oxygen, what colour is the flame?

A green

B orange

C red

D white

53 Which statement about the oxides and hydroxides of the Group II elements Mg, Ca, Sr and Ba is correct?

A Each of the oxides reacts readily with water to form a solution of pH 12 or above.

B Magnesium oxide is used as a furnace lining because it has a giant molecular structure and hence a high melting point.

C The hydroxides are produced directly by the thermal decomposition of the corresponding nitrates.

D The solubility of the hydroxides increases from $Mg(OH)_2$ to $Ba(OH)_2$.

54 **X** is an s-block element. **X** forms an insoluble carbonate with the formula $\mathbf{X}CO_3$. **X** forms a hydroxide that is more soluble than strontium hydroxide.

What could be the identity of **X**?

A barium

B calcium

C cobalt

D rubidium

55 Which substance will **not** be a product of the thermal decomposition of hydrated magnesium nitrate?

A dinitrogen monoxide

B magnesium oxide

C oxygen

D steam

56 When strontium is burnt in oxygen, what colour is the flame?

A green

B red

C white

D yellow

57 Sir Humphrey Davy discovered boron, calcium, magnesium and sodium.

Which of these elements has the **second** smallest atomic radius in its group and the **third** lowest first ionisation energy in its period?

A boron

B calcium

C magnesium

D sodium

58 A sample of potassium oxide, K_2O, is dissolved in $250\,cm^3$ of distilled water. $25.0\,cm^3$ of this solution is titrated against sulfuric acid of concentration $2.00\,mol\,dm^{-3}$. $15.0\,cm^3$ of this sulfuric acid is needed for complete neutralisation.

Which mass of potassium oxide was originally dissolved in $250\,cm^3$ of distilled water?

A 2.83 g **B** 28.3 g **C** 47.1 g **D** 56.6 g

59 Pieces of calcium and magnesium are reacted separately with cold water.

Pieces of calcium and magnesium are also burned separately with oxygen.

Which row correctly describes the observations from these reactions?

	calcium with cold water	magnesium with cold water	calcium with oxygen	magnesium with oxygen
A	rapid bubbling	slow bubbling	red flame	white flame
B	rapid bubbling	slow bubbling	white flame	red flame
C	slow bubbling	rapid bubbling	red flame	white flame
D	slow bubbling	rapid bubbling	white flame	red flame

60 **X** is a mixture of two compounds of Group **II** elements.

X can undergo thermal decomposition to produce a white solid and only two gaseous products. One of the gaseous products relights a glowing splint.

What could be the components of mixture **X**?

A $MgCl_2$ and $CaCO_3$

B $MgCO_3$ and $Ca(NO_3)_2$

C $Mg(NO_3)_2$ and $Ca(NO_3)_2$

D MgO and $CaSO_4$

61 The three minerals below are obtained from mines around the world. Each one behaves as a mixture of two carbonate compounds. They can be used as fire retardants because they decompose in the heat, producing CO_2. This gas smothers the fire.

barytocite	$BaCa(CO_3)_2$
dolomite	$CaMg(CO_3)_2$
huntite	$Mg_3Ca(CO_3)_4$

What is the order of effectiveness as fire retardant, from best to worst?

	best →		worst
A	dolomite	barytocite	huntite
B	dolomite	huntite	barytocite
C	huntite	barytocite	dolomite
D	huntite	dolomite	barytocite

62 Which mass of solid residue will be obtained from the thermal decomposition of 4.10 g of anhydrous calcium nitrate?

A 0.70 g **B** 1.00 g **C** 1.40 g **D** 2.25 g

63 The shell of a chicken's egg makes up 5% of the mass of an average egg. An average egg has a mass of 50 g.

Assume the egg shell is pure calcium carbonate.

How many complete chicken's egg shells would be needed to neutralise 50 cm^3 of 2.0 mol dm^{-3} ethanoic acid?

A 1 **B** 2 **C** 3 **D** 4

64 When 3.00 g of an anhydrous nitrate of a Group II metal is decomposed, 1.53 g of gas is produced.

What is the nitrate compound?

A beryllium nitrate

B calcium nitrate

C magnesium nitrate

D strontium nitrate

65 What happens when a piece of magnesium ribbon is placed in cold water?

A A vigorous effervescence occurs.

B Bubbles of gas form slowly on the magnesium.

C The magnesium floats on the surface of the water and reacts quickly.

D The magnesium glows and a white solid is produced.

66 Compound **X** releases carbon dioxide gas and forms a white solid, **Y**, when it is heated. Neither **X** nor **Y** are soluble in water. Compound **Y** is used as a refractory kiln lining.

What is compound **X**?

A $CaCO_3$ **B** CaO **C** $MgCO_3$ **D** MgO

67 1.00 g of a metallic element reacts completely with 300 cm^3 of oxygen at 298 K and 1 atm pressure to form an oxide which contains O^{2-} ions.

The volume of one mole of gas at this temperature and pressure is 24.0 dm^3.

What could be the identity of the metal?

A calcium

B magnesium

C potassium

D sodium

68 Due to their similar ionic radii, the reactions of lithium and magnesium and their corresponding compounds are very similar.

Which statement about the reactions of lithium or its compounds can be predicted from this statement?

A Lithium burns very slowly in oxygen.

B Lithium carbonate decomposes on heating in a blue Bunsen burner flame, forming lithium oxide and carbon dioxide.

C Lithium nitrate decomposes on heating, forming lithium nitrite, $LiNO_2$, and oxygen.

D Lithium reacts very violently with cold water, liberating hydrogen.

69 **X** is a Group **II** metal. The carbonate of **X** decomposes when heated in a Bunsen flame to give carbon dioxide and a white solid residue as the only products. This white solid residue is sparingly soluble in water. Even when large amounts of the solid residue are added to water the pH of the saturated solution is less than that of limewater.

What could be the identity of **X**?

A magnesium

B calcium

C strontium

D barium

70 Rat poison needs to be insoluble in rain water but soluble at the low pH of stomach contents.

What is a suitable barium compound to use for rat poison?

A barium carbonate

B barium chloride

C barium hydroxide

D barium sulfate

71 Calcium forms an ionic compound with carbon, called calcium carbide. The oxidation number of carbon in calcium carbide is –1.

Calcium carbide is readily hydrolysed by water giving two products only.

What could be the formulae of calcium carbide and the two products of hydrolysis?

	calcium carbide	products
A	Ca_2C	CaO and C_2H_4
B	Ca_2C	$Ca(OH)_2$ and C_2H_2
C	CaC_2	CaO and C_2H_4
D	CaC_2	$Ca(OH)_2$ and C_2H_2

72 Which property decreases on descending Group II?

A radius of the cation, M^{2+}

B reactivity of the element with water

C shielding of outermost electrons

D the ease of thermal decomposition of the carbonates, MCO_3

73 The reaction between aluminium powder and anhydrous barium nitrate is used as the propellant in some fireworks. The metal oxides and nitrogen are the only products.

Which volume of nitrogen, measured under room conditions, is produced when 0.783 g of anhydrous barium nitrate reacts with an excess of aluminium?

A 46.8 cm^3 **B** 72.0 cm^3 **C** 93.6 cm^3 **D** 144 cm^3

74 Which row correctly identifies the uses of some of the compounds of Group II metals?

	used as a refractory lining in kilns	used in agriculture to increase the pH of a soil
A	CaO	$Ca(OH)_2$
B	CaO	$Mg(OH)_2$
C	MgO	$Ca(OH)_2$
D	MgO	$Mg(OH)_2$

75 River water in a chalky agricultural area contains Ca^{2+}, Mg^{2+}, CO_3^{2-}, HCO_3^-, Cl^- and NO_3^- ions. In a water treatment plant, such water is treated by adding a calculated quantity of calcium hydroxide.

What will be precipitated from the river water following the addition of calcium hydroxide?

A $CaCl_2$ **B** $CaCO_3$ **C** $Ca(NO_3)_2$ **D** $Mg(NO_3)_2$

76 X and Y are both Group 2 metals.

X and Y both form hydroxide compounds, but $X(OH)_2$ is more soluble in water than $Y(OH)_2$.

If a piece of metal Y is put into cold water a very slow reaction occurs, and only a very few, small hydrogen bubbles can be seen.

What could be the identities of X and Y?

	X	Y
A	barium	magnesium
B	barium	strontium
C	calcium	strontium
D	magnesium	calcium

77 In some areas lime, $Ca(OH)_2$, is added to soil to improve crop growth.

Which statement correctly describes a reason why lime improves crop growth?

A Lime acts as a catalyst which speeds up the release of nitrates into the soil.

B Lime is an effective pesticide and protects the plants from damage.

C Lime is used to reduce the acidity of the soil.

D Lime lowers the pH of the soil.

78 A solid, **T**, was placed in an excess of the liquid **U**.

A colourless gas was given off and a white precipitate was seen. The precipitate was not **T**.

What could be the identities of **T** and **U**?

	T	**U**
A	$BaCO_3$	H_2O
B	Ca	dilute H_2SO_4
C	Mg	dilute H_2SO_4
D	$SrCO_3$	dilute HCl

79 Solutions of $Mg(NO_3)_2$ and $Ba(NO_3)_2$ separately undergo a series of reactions.

$$Mg(NO_3)_2(aq) \xrightarrow{\text{sodium carbonate solution}} M \xrightarrow{\text{excess } HCl(aq)} N \xrightarrow{\text{excess } NaOH(aq)} P$$

$$Ba(NO_3)_2(aq) \xrightarrow{\text{sodium carbonate solution}} Q \xrightarrow{\text{excess } HCl(aq)} R \xrightarrow{\text{excess } NaOH(aq)} S$$

M, N and P are magnesium compounds.

Q, R and S are barium compounds.

How many of M, N, P, Q, R and S are white precipitates?

A 2 **B** 3 **C** 4 **D** 5

80 Anhydrous magnesium nitrate, $Mg(NO_3)_2$, will decompose when heated, giving a white solid and a mixture of two gases X and Y.

Y is oxygen.

What is the ratio $\dfrac{\text{mass of X released}}{\text{mass of Y released}}$?

A $\dfrac{1}{0.174}$ **B** $\dfrac{1}{0.267}$ **C** $\dfrac{1}{0.348}$ **D** $\dfrac{1}{3.43}$

81 The table gives information about calcium carbonate and calcium hydroxide.

Which row is correct?

	calcium carbonate is more soluble than calcium hydroxide	calcium hydroxide can be manufactured using calcium carbonate as a starting material
A	no	no
B	no	yes
C	yes	no
D	yes	yes

82 Which row of the table gives correct comparisons between the solubilities of calcium and barium hydroxide and the thermal stabilities of calcium and barium carbonate?

	solubility		thermal stability	
	calcium hydroxide	barium hydroxide	calcium carbonate	barium carbonate
A	higher	lower	higher	lower
B	higher	lower	lower	higher
C	lower	higher	higher	lower
D	lower	higher	lower	higher

83 When heated, magnesium nitrate decomposes.

Which equation for the thermal decomposition of magnesium nitrate is correct?

A $Mg(NO_3)_2 \rightarrow MgO + NO_2 + NO + O_2$

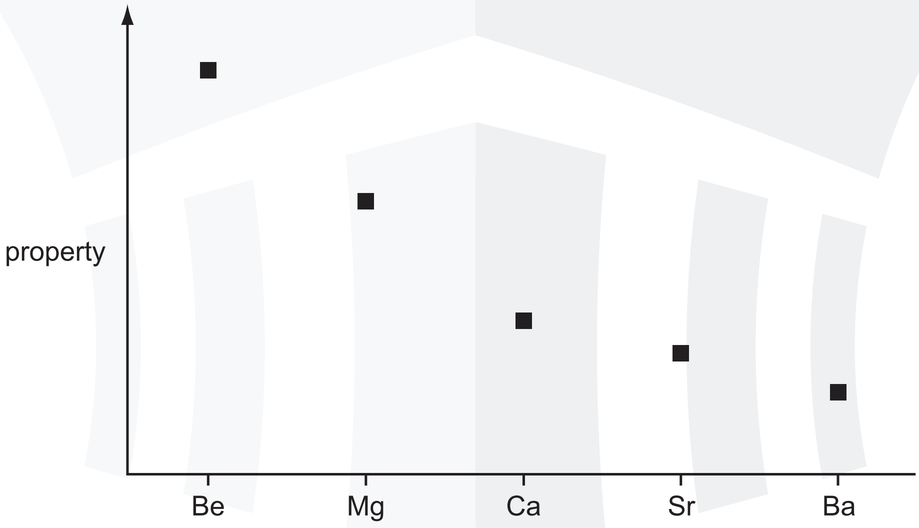
B $2Mg(NO_3)_2 \rightarrow 2MgO + 4NO + 3O_2$

C $2Mg(NO_3)_2 \rightarrow 2MgO + 4NO_2 + O_2$

D $3Mg(NO_3)_2 \rightarrow Mg_2N_3 + MgO + 3NO + 7O_2$

84 *Use of the Data Booklet is relevant to this question.*

The graph represents the variation of a property of the Group II elements.



What is this property?

A ionic radius

B ionisation energy

C neutron / proton ratio

D rate of reaction with water

85 Which property **increases** in value going down Group II?

A electronegativity

B ionic radius

C maximum oxidation number

D second ionisation energy

[W'13 2 Q14]

86 Which property **increases** in value going down Group II?

A electronegativity

B ionic radius

C maximum oxidation number

D second ionisation energy

[W'13 2 Q14]

87 Radium is an element below barium in Group 2 of the Periodic Table.

Which equation shows what happens when solid radium nitrate, $Ra(NO_3)_2$, is heated strongly?

A $Ra(NO_3)_2(s) \rightarrow RaO(s) + N_2O(g) + 2O_2(g)$

B $2Ra(NO_3)_2(s) \rightarrow 2RaO(s) + 2N_2(g) + 5O_2(g)$

C $2Ra(NO_3)_2(s) \rightarrow 2RaO(s) + 4NO_2(g) + O_2(g)$

D $4Ra(NO_3)_2(s) \rightarrow 2Ra_2O(s) + 8NO_2(g) + 3O_2(g)$

[W'17 2 Q14]

88 Exactly 3.705 kg of substance Y are needed to neutralise 100 moles of $HCl(aq)$.

What could be substance Y?

A Ca **B** CaO **C** $Ca(OH)_2$ **D** $CaCO_3$

[W'17 2 Q15]

89 The mineral dolomite is a mixture of magnesium carbonate and calcium carbonate.

An aqueous reagent, X, was added to a small sample of dolomite. Effervescence was seen and a white solid, Y, was formed.

What could be the correct identity of reagent X and solid Y?

	reagent X	solid Y
A	hydrochloric acid	calcium chloride
B	hydrochloric acid	magnesium chloride
C	sulfuric acid	calcium sulfate
D	sulfuric acid	magnesium sulfate

[W'17 3 Q15]

90 Which mass of solid residue is obtained from the thermal decomposition of 4.10 g of anhydrous calcium nitrate?

A 0.70 g **B** 1.00 g **C** 1.40 g **D** 2.25 g

[M'18 2 Q5]

91 When burned, sulfur forms a gaseous product X which can be oxidised to produce a gas Y.

Gas Y reacts with water to produce a product Z.

Which row correctly shows the oxidation states of sulfur in X, Y and Z?

	X	Y	Z
A	–2	+4	+4
B	–2	+4	+6
C	+4	+6	+4
D	+4	+6	+6

[M'18 2 Q18]

92 One molecule of ammonia reacts with one molecule of ethyl methanoate, $HCO_2C_2H_5$, to produce one molecule of methanamide, $HCONH_2$, and only one other molecule, X.

One molecule of methanamide decomposes on heating strongly to produce one molecule of ammonia and only one other molecule, Y.

What could be the identities of X and Y?

	X	Y
A	ethanoic acid	carbon monoxide
B	ethanoic acid	hydrogen cyanide
C	ethanol	carbon monoxide
D	ethanol	hydrogen cyanide

[M'18 2 Q19]

93 Aqueous silver nitrate is added to a solution of potassium iodide.

Aqueous ammonia is then added.

What would be observed?

A a cream precipitate that dissolves on addition of aqueous ammonia

B a cream precipitate that does not dissolve on addition of aqueous ammonia

C a yellow precipitate that dissolves on addition of aqueous ammonia

D a yellow precipitate that does not dissolve on addition of aqueous ammonia

[S'18 1 Q16]

94 In Group 2 of the Periodic Table, the properties of the elements and their compounds show regular change down the group.

Which property shows a **decrease** from magnesium to barium?

A the decomposition temperature of the carbonates

B the decomposition temperature of the nitrates

C the solubility of the hydroxides

D the solubility of the sulfates

[S'18 2 Q16]

SECTION B

For the following questions, the responses A to D should be selected on the basis of

A	B	C	D
1, 2 and **3** are correct	**1** and **2** only are correct	**2** and **3** only are correct	**1** only is correct

1 Which properties would be expected for the Group II element, strontium, or its compounds?

1 When heated in oxygen, strontium does not burn.

2 On being heated, strontium carbonate decomposes to give strontium oxide.

3 When strontium oxide is added to water, the solution is alkaline.

2 Which of the following magnesium compounds lose mass when heated by a bunsen flame?

1 magnesium carbonate

2 magnesium nitrate

3 magnesium oxide

3 A farmer spreads lime on land which has already been treated with a nitrogenous fertiliser.

Which reactions will occur over a period of time?

1 $Ca(OH)_2 + CO_2 \longrightarrow CaCO_3 + H_2O$

2 $Ca(OH)_2 + 2H^+(aq) \longrightarrow Ca^{2+}(aq) + 2H_2O$

3 $Ca(OH)_2 + 2NH_4^+(aq) \longrightarrow Ca^{2+}(aq) + 2NH_3 + 2H_2O$

4 Which magnesium compounds lose mass when heated by a Bunsen flame?

1 magnesium carbonate

2 magnesium nitrate

3 magnesium oxide

5 Which reactions involving calcium and its compounds would produce **two** gaseous products?

1 heating solid anhydrous calcium nitrate

2 heating solid anhydrous calcium carbonate

3 adding calcium metal to water

6 Samples of calcium and barium are separately added to beakers of cold water containing a few drops of litmus solution.

Which observations will be made with **only** the calcium and **not** with the barium?

1 A white suspension appears in the water.

2 The solution turns blue.

3 A gas is evolved.

7 Which oxides react with water to give a solution of pH 10 or higher?

1 CaO

2 Na_2O

3 SrO

8 Which properties would be expected for radium, $_{88}Ra$, or its compounds?

1 Radium carbonate would not decompose at the temperature of a Bunsen flame.

2 Radium hydroxide is very insoluble.

3 Radium does not react with cold water.

9 When a firework is lit a fuel and an oxidising agent react.

In such a firework, magnesium is the fuel and barium nitrate is the oxidising agent.

Which solid products are produced when the firework is lit?

1 BaO

2 MgO

3 $Mg(NO_3)_2$

10 Which properties would be expected from radium, $_{88}Ra$, or its compounds?

1 Radium carbonate decomposes only at a very high temperature.

2 Radium hydroxide is very insoluble.

3 Radium does not react with cold water.

11 Hydroxyapatite, a basic calcium phosphate, $Ca(OH)_2.3Ca_3(PO_4)_2$, is the mineral found in bone.

In older people, calcium ions can be lost from the hydroxyapatite, weakening the bone structure. In such cases, strontium salts are administered to strengthen the bone. The strontium ions replace the lost calcium ions in the hydroxyapatite.

Which statements are correct?

1 Strontium ions are nearly the same size as calcium ions and so may easily replace them in the hydroxyapatite.

2 Strontium hydroxide is less soluble than calcium hydroxide and so will precipitate better in the bone structure.

3 There is ionic, covalent and metallic bonding in hydroxyapatite which gives it strength.

12 The electronic structure of the outer shell of the element radium is $7s^2$.

Which statements will be correct for radium within its group?

1 The element will decompose water, liberating hydrogen.

2 The element will show an oxidation number of +2 in all its compounds.

3 Radium has the highest first ionisation energy.

13 Which statements concerning the Group II elements magnesium, calcium and barium are correct?

1 Their reactivity increases with increasing relative atomic mass.

2 The oxidation number exhibited in their stable compounds is +2.

3 On strong heating, their nitrates give off oxygen only.

14 The rock *dolomite* is a double carbonate of magnesium and calcium, $CaCO_3.MgCO_3$.

When heated strongly, a product called *calcined dolomite* is formed which is used to line some furnaces for the production of metals.

Why is *calcined dolomite* used for this purpose?

1 It is a refractory material.

2 It will absorb acidic impurities in metallurgical processes.

3 It will reduce metallic oxides to metals.

15 Which reactions are redox reactions?

1 $CaBr_2 + 2H_2SO_4 \rightarrow CaSO_4 + Br_2 + SO_2 + 2H_2O$

2 $CaBr_2 + 2H_3PO_4 \rightarrow Ca(H_2PO_4)_2 + 2HBr$

3 $CaBr_2 + 2AgNO_3 \rightarrow Ca(NO_3)_2 + 2AgBr$

16 A farmer spreads lime on land which has already been treated with an ammonium nitrate fertiliser.

Which reactions will occur in the treated soil?

1 $Ca(OH)_2 + 2NH_4^+(aq) \rightarrow Ca^{2+}(aq) + 2NH_3 + 2H_2O$

2 $Ca(OH)_2 + 2H^+(aq) \rightarrow Ca^{2+}(aq) + 2H_2O$

3 $Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$

17 A sample containing 0.40 mol of calcium nitrate was decomposed by heating in a roaring Bunsen burner flame until there was no further decomposition.

What are the products of this reaction?

1 0.40 mol of calcium oxide

2 0.40 mol of nitrogen, $N_2(g)$

3 0.40 mol of oxygen, $O_2(g)$

18 Solids **W**, **X**, **Y** and **Z** are compounds of two different Group II metals. Some of their applications are described below.

Compound **W** is used as a refractory lining material in kilns.

Compound **X** is used as a building material. It can also be heated in a kiln to form compound **Y**. When **Y** is hydrated, it forms compound **Z** which is used agriculturally to treat soils.

Which statements about these compounds are correct?

1 More acid is neutralised by 1 g of **W** than by 1 g of **X**.

2 The metallic element in **W** reacts with water more quickly than the metallic element in **Y**.

3 Adding **Z** to a soil decreases the pH of the soil.

19 Solids **W**, **X**, **Y** and **Z** are compounds of two different Group II metals. Some of their applications are described below.

Compound **W** is used as a refractory lining material in kilns.

Compound **X** is used as a building material. It can also be heated in a kiln to form compound **Y**. When **Y** is hydrated, it forms compound **Z** which is used agriculturally to treat soils.

Which statements about these compounds are correct?

1 Adding **W** to water has less effect on pH than adding **Y**.

2 Adding **Z** to soil increases the pH of the soil.

3 The metallic element in **Y** reacts with cold water more quickly than the metallic element in **W**.

20 Solids **W**, **X**, **Y** and **Z** are compounds of two different Group II metals. Some of their applications are described below.

Compound **W** is used as a refractory lining material in kilns.

Compound **X** is used as a building material. It can also be heated in a kiln to form compound **Y**. When **Y** is hydrated, it forms compound **Z** which is used agriculturally to treat soils.

Which statements about these compounds are correct?

1 More acid is neutralised by 2.0 g of **X** than by 2.0 g of **W**.

2 The M_r of **X** is greater than the M_r of **Y** by 44.0.

3 The metallic element in **Y** reacts with cold water more quickly then the metallic element in **W**.

21 The Group II metals have higher melting points than the Group I metals.

Which factors could contribute towards the higher melting points?

1 There are smaller interatomic distances in the metallic lattices of the Group II metals.

2 More electrons are available from each Group II metal atom for bonding the atom into the metallic lattice.

3 Group II metals have a higher first ionisation energy than the corresponding Group I metal.

22 Which statements about calcium oxide are correct?

1 It reacts with cold water.

2 It is produced when calcium nitrate is heated.

3 It can be reduced by heating with magnesium.

23 **X** is a Group II metal. It forms a sulfate which is more soluble than barium sulfate. It forms a hydroxide which is more soluble than calcium hydroxide.

What could be the identity of **X**?

1 strontium

2 magnesium

3 beryllium

24 Which statements concerning the Group II elements magnesium, calcium and barium are correct?

1 Their reactivity increases with increasing relative atomic mass.

2 The oxidation number exhibited in their stable compounds is +2.

3 On strong heating, their nitrates give off oxygen only.

25 Which statements about calcium and strontium compounds are correct?

1 When calcium oxide and strontium oxide are added to water they both produce alkalis.

2 Calcium hydroxide is more soluble than strontium hydroxide.

3 Calcium sulfate is less soluble than strontium sulfate.

26 Which statements about the elements barium and calcium and their compounds are correct?

1 Barium nitrate decomposes at a higher temperature than calcium nitrate.

2 Barium hydroxide is more soluble in water than is calcium hydroxide.

3 Calcium is more reactive with water than is barium.

[W'17 3 Q35]

27 In which pairs do both species have the same number of electrons?

1 ^{35}Cl and ^{37}Cl

2 $^{35}Cl^-$ and ^{40}Ar

3 ^{40}Ar and $^{40}K^+$

[W'18 2 Q32]

GROUP 2 WS 2

1 This question is about the elements in Group II of the Periodic Table, magnesium to barium.

(a) Complete the table below to show the electronic configuration of calcium atoms and of strontium ions, Sr^{2+}.

	1s	2s	2p	3s	3p	3d	4s	4p	4d
Ca	2	2	6						
Sr^{2+}	2	2	6						

[2]

(b) Explain the following observations.

(i) The atomic radii of Group II elements increase down the Group.

....

....

(ii) The strontium ion is smaller than the strontium atom.

....

....

(iii) The first ionisation energies of the elements of Group II decrease with increasing proton number.

....

....

....

....

[4]

(c) Samples of magnesium and calcium are placed separately in cold water and left for some time. In **each case**, describe what you would see and write a balanced equation for each reaction.

(i) magnesium

observation ....

....

equation ....

(ii) calcium

observation ....

....

equation ....
[6]

(d) Strontium nitrate, $Sr(NO_3)_2$ undergoes thermal decomposition.

(i) State one observation you would make during this reaction.

....

....

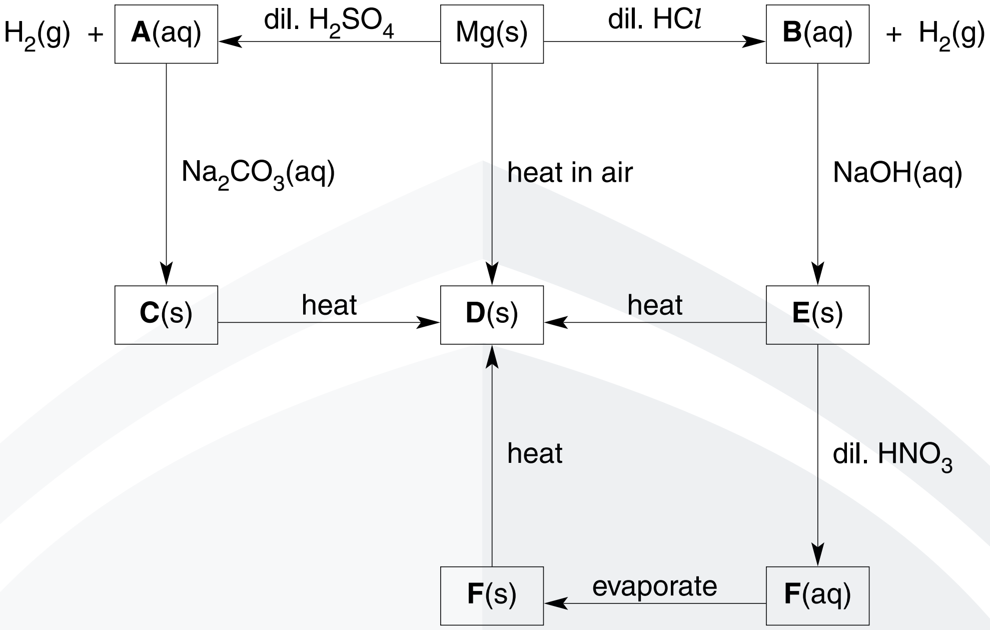
(ii) Write a balanced equation for this reaction.

....
[4]

2 Magnesium is the eighth most common element in the Earth's crust.

The metal is widely used in alloys which are light and strong.

Some reactions of magnesium and its compounds are shown in the reaction scheme below.



(a) Identify, by name or formula, compounds **A** to **F**.

A ....

B ....

C ....

D ....

E ....

F ....

[6]

(b) **(i)** Construct balanced equations for the following reactions.

magnesium to compound **A**

....

compound **C** to compound **D**

....

compound **F** to compound **D**

....

(ii) Suggest a balanced equation for the effect of heat on compound **E**.

.... [4]

3 Radium was discovered in the ore pitchblende by Marie and Pierre Curie in 1898, and the metal was first isolated by them in 1910.

The metal was obtained by first reacting the radium present in the pitchblende to form insoluble radium sulfate which was converted into aqueous radium bromide. This solution was then electrolysed using a mercury cathode and a carbon anode.

(a) Radium has chemical reactions that are typical of Group II metals and forms ionic compounds.

(i) What is the characteristic feature of the electronic configurations of all Group II metals?

....

(ii) Radium sulfate is extremely insoluble. From your knowledge of the simple salts of Group II metals, suggest another very insoluble radium salt.

....

[2]

(c) **(i)** Describe what you would see when magnesium reacts with

cold water, ...

....

steam. ....

....

(ii) Write an equation for the reaction with steam.

....

[5]

(d) Radium reacts vigorously when added to water.

(i) Write an equation, with state symbols, for this reaction.

....

(ii) State **two** observations that could be made during this reaction.

....

....

(iii) Suggest the approximate pH of the resulting solution.

………………

(iv) Will the reaction be more or less vigorous than the reaction of barium with water?

Explain your answer.

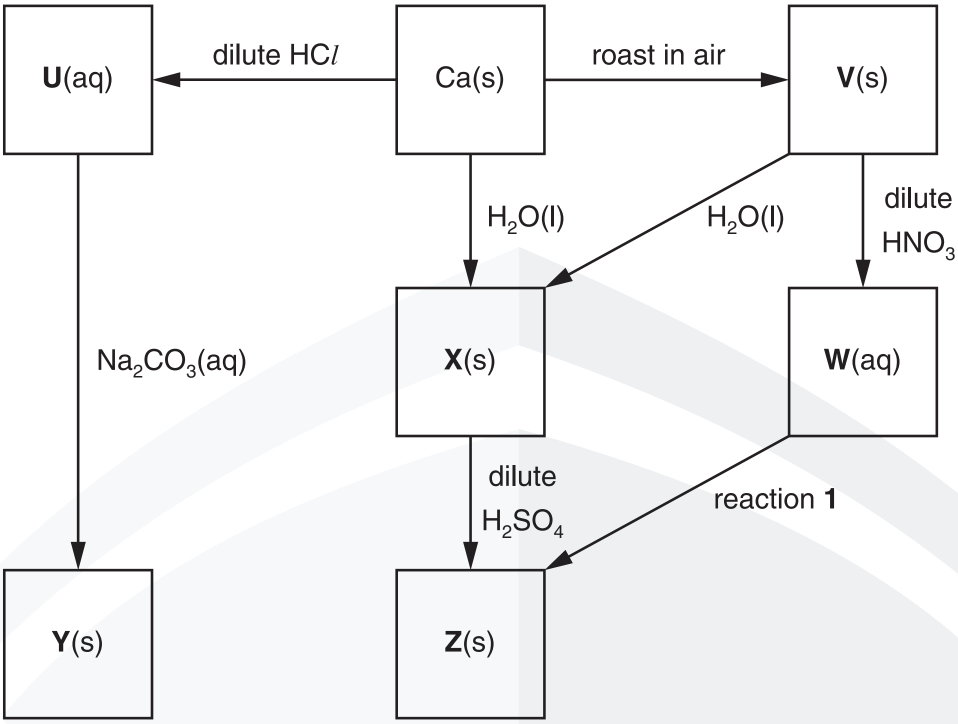
....

....

[6]

4 Calcium is the fifth most common element in the Earth's crust.
Calcium compounds occur in bones and teeth and also in many minerals.

Some reactions of calcium and its compounds are shown in the reaction scheme below.



(a) State the formula of **each** of the calcium compounds **U** to **Y**.

U ....

V ....

W ....

X ....

Y .... [5]

(b) Compound **Y** may be converted into compound **V**.
Outline how this reaction would be carried out in a school or college laboratory using a small sample of **Y**.

....

....[1]

(c) **(i)** Construct balanced equations for the following reactions.

calcium to compound **U**

....

compound **V** to compound **W**

....

compound **U** to compound **Y**

....

(ii) Construct a balanced equation for the effect of heat on solid compound **W**.

....
[4]

(d) Suggest the formula of an aqueous reagent, other than an acid, for reaction **1**.

.... [1]

(e) What would be observed when **each** of the following reactions is carried out in a test-tube?

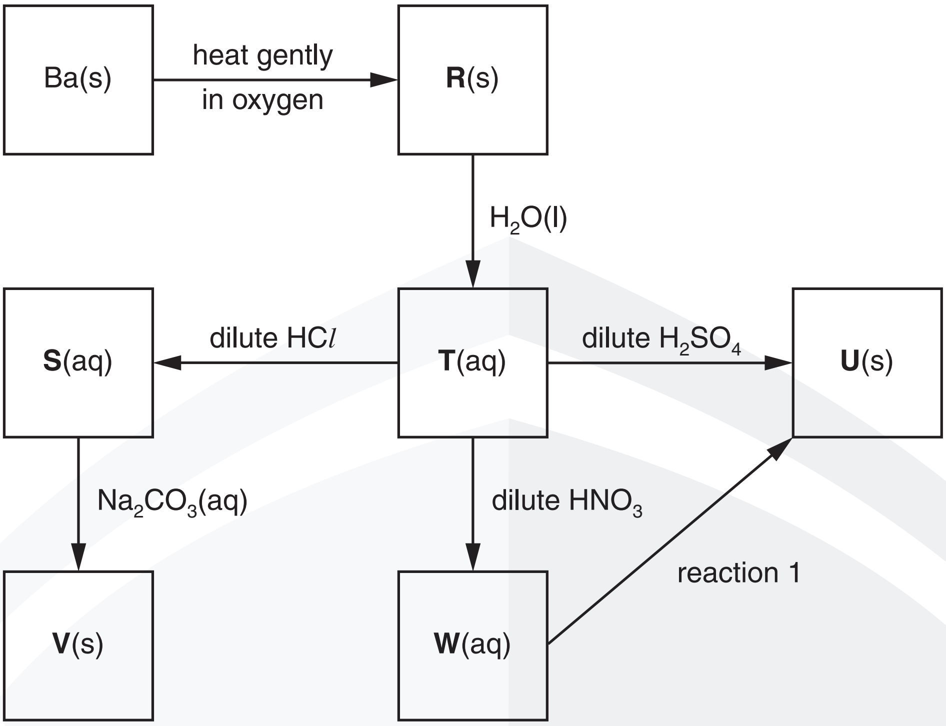
the formation of **X** from Ca(s)

....

the formation of **X** from **V**

....
[2]

5 Barium, proton number 56, is a Group II element which occurs in nature as the carbonate or sulfate.
The element was first isolated by Sir Humphry Davy in 1808.
Some reactions of barium and its compounds are shown in the reaction scheme below.



(a) State the formula of **each** of the barium compounds **R** to **W**.

R .... **S** ....

T.................................. **U** ....

V .... **W** .... [6]

(b) **(i)** Write balanced equations for the following reactions.

compound **T** to compound **W**

....

the roasting of **V** in air

....

(ii) Suggest a gaseous reagent for the conversion of **T** into **V** and write a balanced equation for the reaction.

reagent ....

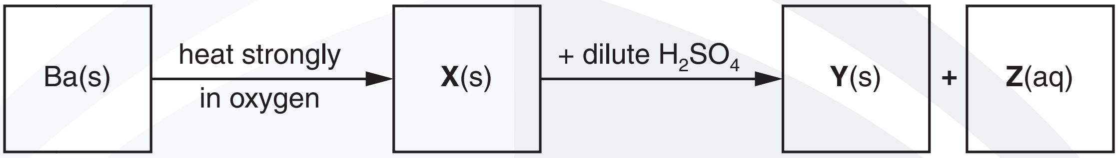
equation ....
[4]

(c) Suggest the formula of an aqueous reagent, other than an acid, for reaction 1.

……………………….................. [1]

When barium is heated strongly in oxygen, an oxide **X** is formed.
The oxide **X** contains 18.9% of oxygen by mass.

The oxide **X** reacts with dilute sulfuric acid in a 1:1 ratio.
Two products, one insoluble and one soluble, are formed.



(d) **(i)** Calculate the empirical formula of **X**.

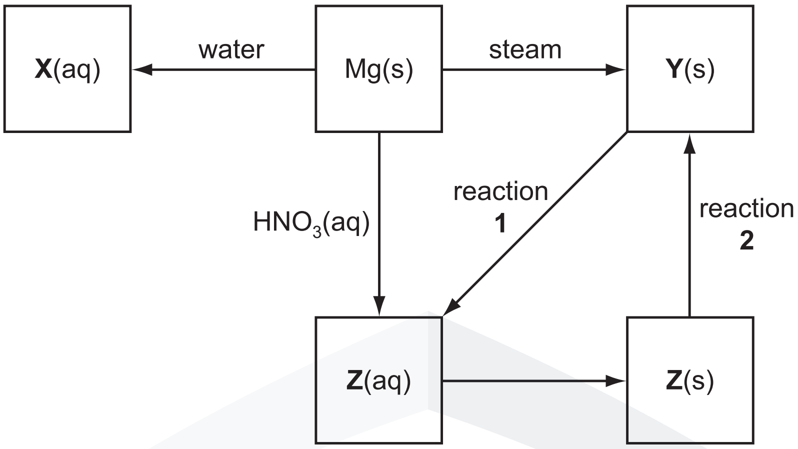
(ii) Suggest the identity of the solid **Y**.

....

(iii) Use your answers to **(i)** and **(ii)** to construct an equation for the reaction of **X** with H_2SO_4.

.... [4]

6 **(d)** Some reactions involving magnesium and its compounds are shown in the reaction scheme below.



(i) Give the **formulae** of the compounds **X**, **Y** and **Z**.

X ....

Y ....

Z ....

[3]

(ii) Name the reagent needed to convert **Y**(s) into **Z**(aq) in reaction **1** and write an equation for the reaction.

reagent ....

equation ....

[2]

(iii) How would you convert a sample of **Z**(s) into **Y**(s) in reaction **2**?

.... [1]

(iv) Give equations for the conversions of Mg into **X**, and **Z**(s) into **Y**.

Mg to **X** ....

Z to **Y** ....

[2]

7 The elements in Group II, and their compounds, show a variety of trends in their properties.

(a) Magnesium, calcium and barium all react with cold water to form hydroxides.

(i) Describe and explain the trend in reactivity of these three elements with cold water.

....

....

....

....

.... [3]

(ii) Give the equation for the reaction of magnesium with cold water.

.... [1]

(iii) Suggest why the water eventually turns cloudy during the reaction of magnesium with cold water.

....

.... [1]

(iv) Suggest the equation for the reaction of hot magnesium with steam.

.... [1]

(b) The oxides of magnesium, calcium and barium all react with dilute nitric acid to form nitrates.

(i) Give the equation for the reaction of magnesium oxide with nitric acid.

.... [1]

(ii) State the trend in thermal stability of the nitrates of Group II.

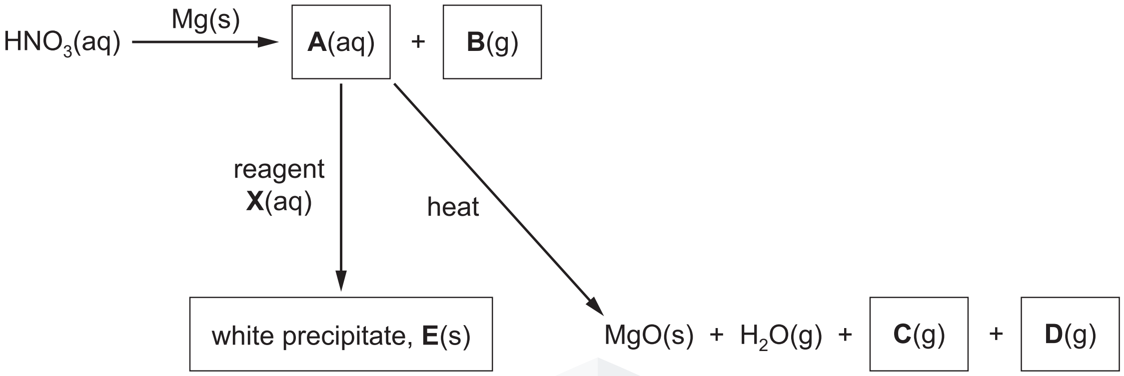
....

.... [1]

(iii) Give the equation for the thermal decomposition of magnesium nitrate.

.... [1]

8 The flow chart below shows a series of reactions.



(i) Give the **formula** of each of the compounds **A** to **D**.

A .... **B** ....

C .... **D** ....

[4]

(ii) **E** reacts with dilute aqueous acid to produce a gas that turns limewater cloudy.

Suggest the identity of reagent **X**.

.... [1]

9 The elements in Group 2, and their compounds, show many similarities and trends in their properties.

(a) Magnesium, calcium, strontium and barium all react with cold water.

(i) Describe what you would **see** when some calcium is added to cold water.

....

....

.... [3]

(ii) Write an equation for the reaction taking place in **(i)**.

.... [1]

(iii) Describe how the reaction of barium with cold water would differ from the reaction of calcium in **(i)** in terms of what you would **see**.

....

.... [1]

(b) Magnesium oxide is used to manufacture heat-resistant bricks for furnace linings in the steel-making industry.

State and explain the property of magnesium oxide that makes it suitable for this use.

....

....

....

.... [2]

(ii) Suggest a reason why magnesium oxide cannot be used as a lining for any furnaces containing acidic materials.

....

.... [1]

(c) The nitrates and carbonates of the Group 2 elements, from magnesium to barium, decompose when heated.

(i) State the trend in the temperature of thermal decomposition of these Group 2 nitrates and carbonates.

....

.... [1]

(ii) Give the equation for the thermal decomposition of magnesium carbonate.

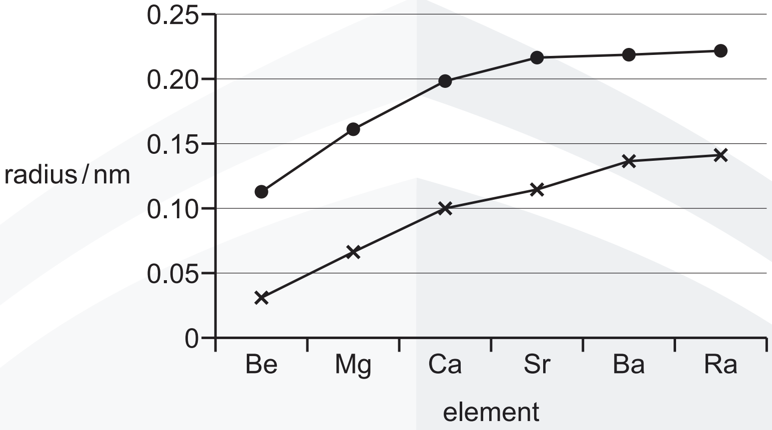
.... [1]

(iii) Give the equation for the thermal decomposition of calcium nitrate.

.... [1]

10 The elements in Group 2 and their compounds show various trends in their physical and chemical properties.

(a) The graph below shows the radius values of the atoms and 2+ ions of the elements in Group 2.



(i) Explain why both lines show a steady increase in the values of the radii down the group.

....

....

.... [2]

(ii) State and explain which line represents the atomic radii and which represents the ionic radii.

....

....

.... [2]

(b) **L** is a salt of a Group 2 element **M**.
When **L** is heated strongly a brown gas is observed and a white solid remains.

The white solid dissolves in water to form a colourless solution of the metal hydroxide **M**$(OH)_2$.

Addition of dilute sulfuric acid to this colourless solution produces a dense white precipitate.

(i) Identify the anion in salt **L**.

.... [1]

(ii) Identify the element **M** and write an **ionic** equation for the formation of the white precipitate with sulfuric acid.

M = ...

equation .... [1]

(iii) Give the formula of salt **L** and use it to write an equation for the thermal decomposition of salt **L**.

formula of salt **L** ....

equation .... [2]

(c) Calcium carbonate and calcium hydroxide can both be used in agriculture to neutralise acidic soils.

(i) Write **ionic** equations for the neutralisation of acid by each of calcium hydroxide and calcium carbonate.

calcium hydroxide ....

calcium carbonate .... [2]

(ii) Suggest and explain why calcium carbonate is a better choice than calcium hydroxide for this purpose in areas of high rainfall.

....

....

.... [2]

(d) Magnesium reacts with both cold water and steam.

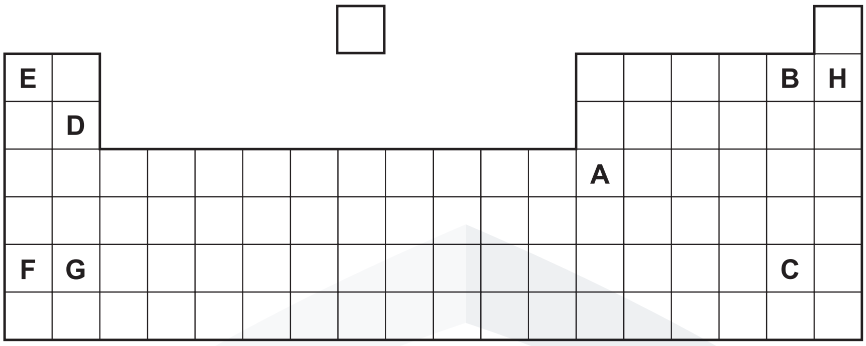
Give the formula of the magnesium-containing product of each of these reactions.

with cold water ....

with steam .... [2]

11 The properties of elements and their compounds show similarities, differences and trends depending on the positions of the elements in the Periodic Table.

(a) The positions of some elements are indicated. The letters used are **not** the symbols of the elements.



From the elements labelled, give the letter for;

(i) the element that forms an amphoteric oxide, ... [1]

(ii) the element with the highest first ionisation energy, ... [1]

(iii) the element that forms a soluble hydroxide and an insoluble sulfate, ... [1]

(iv) the most volatile element in a group that contains elements in all three states of matter at room temperature and pressure,

.... [1]

(v) the element that forms the largest cation. .... [1]

(b) The elements in Group 2 all react with oxygen and with water.

(i) State and explain the conditions needed for magnesium to react with oxygen.

....

.... [2]

(ii) State what would be seen during the reaction in **(b)(i)**.

....

.... [1]

(iii) Write an equation for the reaction of magnesium with cold water.
Include state symbols.

.... [2]

(c) The carbonates and nitrates of the elements in Group 2 can all be decomposed by heating.

(i) Write an equation for the thermal decomposition of magnesium nitrate.

.... [1]

(ii) The thermal decomposition of calcium carbonate forms a solid product that is industrially important. This solid product reacts with water to form a compound commonly known as slaked lime.

Write equations for the thermal decomposition of calcium carbonate and the reaction of the solid product to form slaked lime.

thermal decomposition ....

formation of slaked lime ....

[2]

(d) Calcium carbonate and calcium hydroxide both have an important use in agriculture.

(i) Describe this use and explain what makes these two compounds suitable for it.

....

....

.... [2]

(ii) Write an ionic equation to illustrate this use of calcium carbonate.

.... [1]

[S'17 1 Q3]

12 The elements in Group 2 show trends in their properties that are typical of metals.
The elements in Group 17 show trends in their properties that are typical of non-metals.

(a) State and explain the trend in ionisation energy down Group 2.

....

....

.... [2]

(ii) The melting point decreases down Group 2.

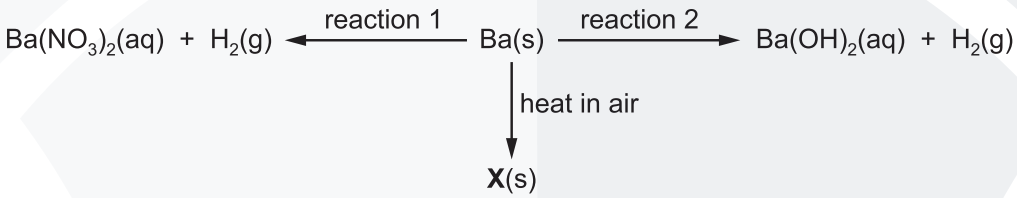
Explain this trend.

....

....

.... [2]

(c) Some reactions based on the Group 2 metal barium, Ba, are shown.



(i) State the reagent needed for each of reactions 1 and 2.

reaction 1 ....

reaction 2 ....
[2]

(ii) Name **X** and write an equation for its formation.

name ....

equation ....
[2]

(iii) The $Ba(NO_3)_2(aq)$, produced by reaction 1, is heated to dryness. The anhydrous solid is then heated strongly and decomposes. Barium oxide is produced, together with two other products.

Identify the **two** other products of this decomposition reaction and state what would be observed.

....

.... [2]

(iv) State what would be observed when excess $MgSO_4(aq)$ is added to the $Ba(OH)_2(aq)$ produced in reaction 2. Explain your answer.

....

....

....

.... [3]

[S'17 3 Q3]

13 Calcium and its compounds have a large variety of applications.

(a) Calcium metal reacts readily with most acids.

(i) Write an equation for the reaction of calcium with dilute nitric acid. State symbols are **not** required.

.... [1]

(ii) When calcium metal is placed in dilute sulfuric acid, it reacts vigorously at first.

After a short time, a crust of calcium sulfate forms on the calcium metal and the reaction stops. Some of the calcium metal and dilute sulfuric acid remain unreacted.

Suggest an explanation for these observations.

....

....

....

.... [2]

(c) Calcium chlorate(I), $Ca(ClO)_2$, is used as an alternative to sodium chlorate(I), NaClO, in some household products.

(i) Suggest a use for calcium chlorate(I).

.... [1]

(ii) The chlorate(I) ion is formed when cold aqueous sodium hydroxide reacts with chlorine.

Write an **ionic** equation for this reaction. State symbols are **not** required.

.... [1]

(iii) The chlorate(I) ion is unstable and decomposes when heated as shown.

Deduce the oxidation number of chlorine in each species. Complete the boxes.

$$3ClO^- \rightarrow 2Cl^- + ClO_3^-$$

oxidation number of chlorine: $\boxed{+1}$ $\boxed{}$ $\boxed{}$

[1]

(iv) In terms of electron transfer, state what happens to the chlorine in the reaction in **(iii)**.

.... [1]

[S'17 3 Q2]

14 Sodium and magnesium are the first two elements in the third period.

(a) Sodium and magnesium both react with cold water to produce the same type of product in solution. With sodium the solution is clear but with magnesium it appears cloudy.

(i) Write an equation for the reaction of magnesium with cold water.

.... [1]

(ii) Suggest why the solution is cloudy after the reaction of magnesium with cold water.

....

.... [1]

(b) Group 2 elements, including magnesium, react with oxygen and with dilute acids. There are trends in both the physical and chemical properties of the elements and their compounds down the group. Reactivity generally increases from Mg to Ba.

(i) Explain why there is a general increase in reactivity from Mg to Ba.

....

....

.... [2]

(ii) Give **two** observations for the reaction of magnesium with oxygen. Write an equation for this reaction. Include state symbols.

....

....

equation ....

[3]

(iii) Write an equation for the reaction of magnesium with sulfuric acid.

.... [1]

(iv) Suggest why there is a general **decrease** in the melting points of the elements down Group 2.

....

....

....

....

.... [3]

[S'18 2 Q1]

NOTES

11 Group 17

The physical and chemical properties of the elements of Group 17 (the halogens) are introduced in this topic.

11.1 Physical properties of the Group 17 elements ☐

11.2 The chemical properties of the elements and their hydrides ☐

11.3 Some reactions of the halide ions ☐

11.4 The reactions of chlorine with aqueous sodium hydroxide ☐

11.5 Some important uses of the halogens and of halogen compounds ☐

GROUP 17

11 Group 17

The physical and chemical properties of the elements of Group 17 (the halogens) are introduced in this topic.

	Learning outcomes Candidates should be able to:
11.1 Physical properties of the Group 17 elements	a) describe the colours and the trend in volatility of chlorine, bromine and iodine b) interpret the volatility of the elements in terms of van der Waals' forces
11.2 The chemical properties of the elements and their hydrides	a) describe the relative reactivity of the elements as oxidising agents (see also Section 6.3(f)) b) describe and explain the reactions of the elements with hydrogen c) (i) describe and explain the relative thermal stabilities of the hydrides (ii) interpret these relative stabilities in terms of bond energies
11.3 Some reactions of the halide ions	a) describe and explain the reactions of halide ions with: (i) aqueous silver ions followed by aqueous ammonia (ii) concentrated sulfuric acid
11.4 The reactions of chlorine with aqueous sodium hydroxide	a) describe and interpret, in terms of changes of oxidation number, the reaction of chlorine with cold and with hot aqueous sodium hydroxide
11.5 Some important uses of the halogens and of halogen compounds	a) explain the use of chlorine in water purification b) state the industrial importance and environmental significance of the halogens and their compounds (e.g. for bleaches, PVC, halogenated hydrocarbons as solvents, refrigerants and in aerosols. See also Section 16.2)

INTRODUCTION

The Halogens consist of the following elements:

- Fluorine (F), Chlorine (Cl), Bromine (Br), Iodine (I) and Astatine (At).

All the halogens consist of diatomic molecules, X_2, linked by a single covalent bond. They are all volatile.

Fluorine (F)	$1s^2\ 2s^2\ 2p^5$
Chlorine (Cl)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$
Bromine (Br)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^5$
Iodine (I)	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}\ 4s^2\ 4p^6\ ^4d^{10}\ 5s^2\ 5p^5$

PHYSICAL PROPERTIES

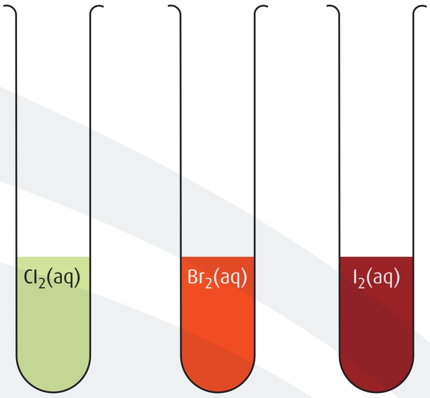
Halogen	Chlorine	Bromine	Iodine
Melting point/°C	–101	–7	114
Boiling point/°C	–34	59	184
Appearance at room temperature	yellow-green gas	red-brown liquid	nearly black solid
$E(X—X)$/kJ mol^{-1}	242	193	151
$E(H—X)$/kJ mol^{-1}	431	366	299

The van der Waals' forces between the molecules arise from the attractions between temporary and induced dipoles. The greater the number of electrons in a molecule, the more polarisable it is and the greater the possibility for temporary induced dipoles. This explains why the boiling temperatures increase down the group of halogens.

COLOURS OF SOLUTIONS

Halogens in a non-polar organic solvent.

Chlorine solution is pale yellow-green, bromine solution is orange & iodine solution is red brown.



HALOGENS AS OXIDANTS

The halogens are powerful oxidising agents. Apart from fluorine, chlorine is the strongest oxidising agent among these elements.

The tendency to form negative ions decreases down the group. The shielding effect of the inner electrons means that the effective nuclear charge that attracts the outer electron is 7+ for all of the atoms.

Down the group the outer electrons get further and further away from the same effective nuclear charge and so they are held less strongly and the tendency to gain an extra electron decreases. Hence reactivity of halogens decreases down the group.

HALOGENS WITH HYDROGEN

All the halogens can react with hydrogen, but the conditions for the reaction vary going down the group.

The reactivity with hydrogen decreases down the group because the decreasing strength of the H—X (HCl, HBr, HI) bond makes the reaction less exothermic.

Halogen	Reaction Condition	Equation	H—X Bond Energy
Chlorine	explodes when exposed to ultraviolet light	$H_{2\,(g)} + Cl_{2\,(g)} \longrightarrow 2HCl_{(g)}$	431 kJ/mol
Bromine	slow reaction when heated	$H_{2\,(g)} + Br_{2\,(g)} \longrightarrow 2HBr_{(g)}$	366 kJ/mol
Iodine	reversible reaction when heated	$H_{2\,(g)} + I_{2\,(g)} \rightleftharpoons 2HI_{(g)}$	299 kJ/mol

HYDROGEN HALIDES

The hydrogen halides are all colourless gases with the formula HX, where X stands for Cl, Br or I. The bonds between hydrogen and the halogens are polar. Hydrogen chloride, hydrogen bromide and hydrogen iodide are similar in that they are very soluble in water to form acidic solutions that ionise completely in water.

Halogen	Bond Energy
H—Cl	431 kJ/mol
H—Br	366 kJ/mol
H—I	299 kJ/mol

HYDROGEN HALIDES

The bond energies decrease going down Group 17, making it easier to break the H—X bond. This is because the iodine atom is the largest atom, so the overlap of its outer shell with a hydrogen atom gives a much longer bond length than with the other smaller halogen atoms. The longer the bond, the weaker it is, and the less energy required to break it. Hence HI is less thermally stable than HCl.

Halogen	Bond Energy
H—Cl	431 kJ/mol
H—Br	366 kJ/mol
H—I	299 kJ/mol

HYDROGEN HALIDES

The strength of the H—X bonds determines the thermal stability of the hydrogen halides. This is readily illustrated by the observations that are made when they are heated.

Hydrogen chloride does not decompose on heating. Hydrogen bromide decomposes slightly and the gas turns brown. Hydrogen iodide decomposes almost completely and the gas turns purple.

Halogen	Reaction (if any)	Comment
chlorine	$HCl(g)$	no decomposition
bromine	$2HBr(g) \rightleftharpoons H_2(g) + Br_2(g)$	some decomposition and gas appears slightly brown
iodine	$2HI(g) \rightarrow H_2(g) + I_2(g)$	almost complete decomposition and gas appears deep purple

HALIDES

Halide ions are the ions of the halogen elements in oxidation state −1. They include the chloride, Cl^-, bromide, Br^-, and iodide, I^- , ions.

In Group 17, a more reactive halogen oxidises the ions of a less reactive halogen. And as halogens are oxidising agents, **a halogen that is a more powerful oxidising agent displaces one that is less powerful.** For example, bromine reacts with a solution of an iodide to produce iodine and a bromide. This is because bromine has a stronger tendency to gain electrons and turn into ions than iodine.

$$Br_{2\,(aq)} + 2I^-_{(aq)} \longrightarrow 2Br^-_{(aq)} + I_{2\,(s)}$$

But iodine will not displace bromide or chloride ions as iodine is the weakest of the three oxidising agents.

HALIDE REACTIONS WITH SILVER IONS

Silver nitrate solution can be used to distinguish between halides. Silver halides are insoluble. Adding silver nitrate to a solution of one of these halide ions produces a precipitate.

$$Ag^+_{(aq)} + X^-_{(aq)} \longrightarrow AgX_{(s)}$$

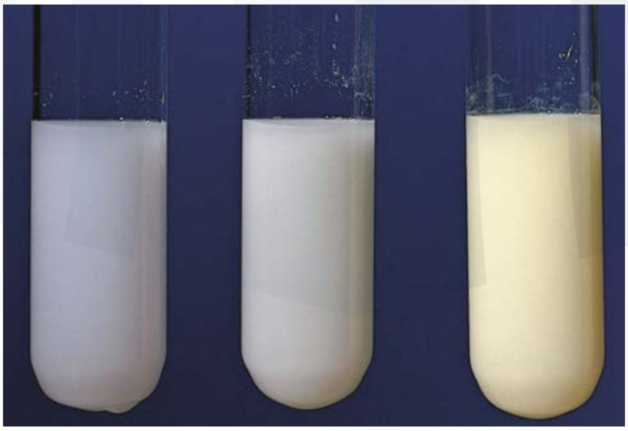
Silver chloride is a white solid. This distinguishes it from silver bromide, which is a creamy colour, and silver iodide, which is a brighter yellow.

The colour changes are not very distinct, but a further test with ammonia helps to distinguish the precipitates. Silver chloride easily dissolves in dilute ammonia solution. Silver bromide dissolves in concentrated ammonia solution, but silver iodide does not dissolve in ammonia solution at all.

HALIDE REACTIONS WITH SILVER IONS

AgCl AgBr AgI

HALIDE REACTIONS WITH SILVER IONS



AgCl AgBr AgI

HALIDE REACTIONS WITH SILVER IONS

Halide	$AgNO_3$ (aq)	dilute NH_3 (aq)	concentrated NH_3 (aq)
Chloride	white ppt.	Soluble	-
Bromide	cream ppt.	Insoluble	Soluble
Iodide	pale yellow ppt.	Insoluble	Insoluble

The precipitates are due to the formation of silver halides.

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \longrightarrow AgCl_{(s)}$$

A diammine complex, $[Ag(NH_3)_2]^+$ is formed, which dissolves the precipitate.

$$AgCl_{(s)} + 2NH_{3\,(aq)} \longrightarrow [Ag(NH_3)_2]^{+}_{(aq)} + Cl^{-}_{(aq)}$$

HALIDES AS REDUCTANTS

Halogen anions (halides) are powerful reducing agents.

This is due to halides requiring to lose only one electron to form halogens.

Down the group, reducing power increases. Hence, Fluoride ion is the weakest reducing agent.

This is due to the fact that down a group, an increased ionic radius means an electron can easily be removed from the halide ion.

REDOX REACTIONS

The forward reaction is reduction of **halogens**, and the backward reaction is the oxidation of **halides**.

Chlorine is completely reduced to **chloride** ions, and **iodide** ions are completely oxidised to **iodine**.

Since oxidation power is based on the tendency to gain electrons, and reduction power is based on the tendency to lose electrons:

a strong oxidising **halogen** has a weaker reducing **halide**, and vice versa.

$$Cl_2 + 2e \rightleftharpoons 2Cl^-$$

$$Br_2 + 2e \rightleftharpoons 2Br^-$$

$$I_2 + 2e \rightleftharpoons 2I^-$$

REACTIONS WITH CONC. SULFURIC ACID

Halides can also be distinguished from each other by reacting them with concentrated sulfuric acid:

- Add concentrated sulfuric acid to a solid halide.
- Hydrogen halides are displaced from their salts, all of which give white fumes in air.
- Down the group halides are strong reducing agents, they can further react to reduce the sulfuric acid into many oxidation states.
- These reactions demonstrate that the strength of the halides as reducing agents is in the order $I^- > Br^- > Cl^-$

REACTIONS WITH CONC. SULFURIC ACID

When concentrated sulfuric acid is added to a solid metal halide, the hydrogen halide is produced.

With sodium chloride, steamy fumes of hydrogen chloride are evolved and sodium hydrogensulfate is formed:

$$NaCl_{(s)} + H_2SO_{4(aq)} \longrightarrow NaHSO_{4(aq)} + HCl_{(g)}$$

On heating the acid gives off two H^+ and forms Na_2SO_4.

With sodium bromide and sodium iodide, similar reactions occur initially:

$$NaBr_{(s)} + H_2SO_{4(aq)} \longrightarrow NaHSO_{4(aq)} + HBr_{(g)}$$

$$NaI_{(s)} + H_2SO_{4(aq)} \longrightarrow NaHSO_{4(aq)} + HI_{(g)}$$

REACTIONS WITH CONC. SULFURIC ACID

Concentrated sulfuric acid is a strong oxidising agent. It is unable to oxidise HCl, but it can oxidise HBr and HI. Hence, only a little HBr is formed when concentrated sulfuric acid is added to sodium bromide, the principal products being bromine and sulfur dioxide

$$NaBr_{(s)} + H_2SO_{4(aq)} \longrightarrow NaHSO_{4(aq)} + HBr_{(g)}$$

$$2HBr_{(g)} + H_2SO_{4(aq)} \longrightarrow 2H_2O_{(l)} + SO_{2(g)} + Br_{2(g)}$$

HI is so easily oxidised that only a trace of it is found. The sulfuric acid is reduced not only to SO_2 (+4), but further to sulfur (0), and even to hydrogen sulphide (—2).

REACTIONS WITH CONC. SULFURIC ACID

Halide	Observations	Product	Oxidation State of S	Reaction type
Chloride	White fumes	HCl	—1	Displacement of halide
Bromide	White fumes	HBr	—1	Displacement of halide
	Brown vapour	Br_2	0	Oxidation of halide
	Colourless gas	SO_2	+4	Reduction of acid
Iodide	White fumes	HI	—1	Displacement of halide
	Purple vapour	I_2	0	Oxidation of halide
	yellow solid	S_8	0	Reduction of sulfur
		H_2S	—2	

DISPROPORTIONATION OF CHLORINE

Chlorine oxoanions form when chlorine reacts with water and alkalis. When chlorine dissolves in water, it reacts reversibly to form a mixture of weak chloric(I) acid and strong hydrochloric acid.

$$Cl_2 + H_2O \rightleftharpoons HCl + HOCl$$

This is an example of a reaction in which the same element both increases and decreases its oxidation number.

In other words, some of the element is oxidised while the rest of the element is reduced. Reactions of this type are called **disproportionation** reactions.

Bromine reacts in a similar way but to a much lesser extent. Iodine is almost insoluble in water and hardly reacts at all.

DISPROPORTIONATION OF CHLORINE

When chlorine dissolves in sodium (or potassium) hydroxide solution at room temperature, it produces chlorate(I) and chloride ions.

$$2NaOH + Cl_2 \longrightarrow NaCl + NaClO + H_2O$$

$$2OH^- + Cl_2 \longrightarrow Cl^- + ClO^- + H_2O$$

Chlorine itself, being an element, has the oxidation number 0. In the chloride ion, it has an oxidation number of −1. In the ClO^- ion, the chlorine has an oxidation number of +1 in order to give an overall charge of −1 on the ion. This is another example of disproportionation.

DISPROPORTIONATION OF CHLORINE

When chlorine is passed into hot, concentrated aqueous sodium hydroxide, it produces sodium chloride and sodium chlorate(V).

$$6NaOH + 3Cl_2 \longrightarrow 5NaCl + NaClO_3 + 3H_2O$$

$$6OH^- + 3Cl_2 \longrightarrow 5Cl^- + ClO_3^- + 3H_2O$$

NOTES

GROUP 17 WS 1

1 Properties of chlorine, iodine and their compounds are compared.

Property **Q** for chlorine is smaller than for iodine.

What is property **Q**?

A oxidising ability of the element

B solubility of the silver halide in $NH_3(aq)$

C strength of van der Waals' forces between the molecules of the element

D thermal stability of the hydrogen halide

2 What happens when chlorine is bubbled through aqueous potassium iodide?

A Chlorine is oxidised to chloride ions.

B Hydrochloric acid is formed.

C Iodide ions are oxidised to iodine.

D Potassium iodide is reduced to iodine.

3 Gaseous nitrogen is less reactive than gaseous fluorine.

What is the reason for this difference in reactivity?

A The boiling point of nitrogen is lower than that of fluorine.

B The relative molecular mass of nitrogen is lower than that of fluorine.

C The atomic radius of nitrogen is greater than that of fluorine.

D The bond strength in the molecule is greater in nitrogen than in fluorine.

4 Which statement is most likely to be true for astatine, which is below iodine in Group VII of the Periodic Table?

A Astatine and aqueous potassium chloride react to form aqueous potassium astatide and chlorine.

B Potassium astatide and hot dilute sulfuric acid react to form white fumes of only hydrogen astatide.

C Silver astatide reacts with dilute aqueous ammonia in excess to form a solution of a soluble complex.

D Sodium astatide and hot concentrated sulfuric acid react to form astatine.

5 Photochromic glass, used for sunglasses, darkens when exposed to bright light and becomes more transparent again when the light is less bright. The depth of colour of the glass is related to the concentration of silver atoms.

The following reactions are involved.

reaction 1 $Ag^+ + Cl^- \rightleftharpoons Ag + Cl$

reaction 2 $Cu^+ + Cl \rightarrow Cu^{2+} + Cl^-$

reaction 3 $Cu^{2+} + Ag \rightarrow Cu^+ + Ag^+$

Which statement about these reactions is correct?

A Cu^+ and Cu^{2+} ions act as catalysts.

B Cu^+ ions act as an oxidising agent in reaction 2.

C Reaction 2 is the one in which light is absorbed.

D Ag^+ ions are oxidised in reaction 1.

6 Chlorine compounds show oxidation states ranging from –1 to +7.

What are the reagent(s) and conditions necessary for the oxidation of elemental chlorine into a compound containing chlorine in the +5 oxidation state?

A $AgNO_3$(aq) followed by NH_3(aq) at room temperature

B concentrated H_2SO_4 at room temperature

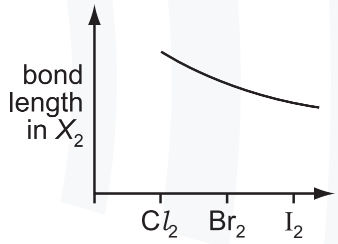
C cold dilute NaOH(aq)

D hot concentrated NaOH(aq)

7 Which graph correctly describes a trend found in the halogen group?

[*X* represents a halogen atom.]

A



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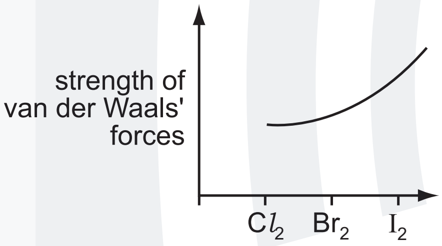

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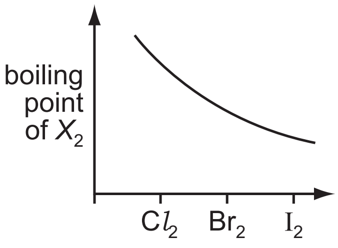
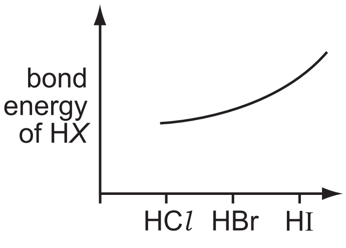
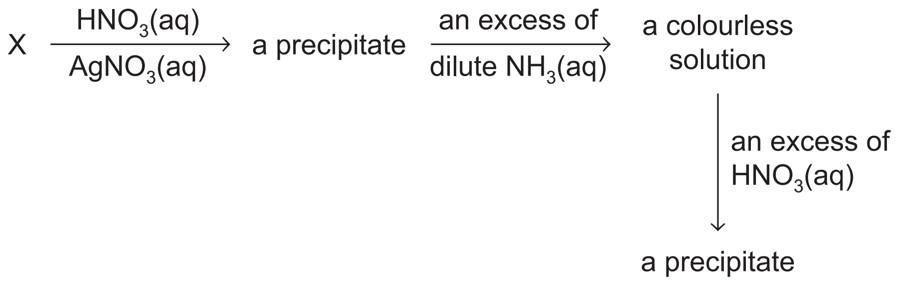


D



8 X is a salt of one of the halogens chlorine, bromine, iodine, or astatine (proton number 85).

The reaction scheme shows a series of reactions using a solution of X as the starting reagent.



What could X be?

- **A** sodium chloride
- **B** sodium bromide
- **C** potassium iodide
- **D** potassium astatide

9 Over half a million tonnes of bromine are manufactured annually and are mainly used for making other compounds. One important use is for agricultural chemicals.

What is another important use for bromine?

- **A** antiseptic agents
- **B** bleaches for textiles and the paper industry
- **C** flame-retardants and fire extinguishers
- **D** water purification

10 How does concentrated sulfuric acid behave when it reacts with sodium chloride?

- **A** as an acid only
- **B** as an acid and oxidising agent
- **C** as an oxidising agent only
- **D** as a reducing agent only

11 Which statement describes the halogens chlorine, bromine and iodine?

- **A** Their bond energies decrease with increasing proton number.
- **B** Their first ionisation energies increase with increasing proton number.
- **C** They are all coloured gases at room temperature.
- **D** They are all good reducing agents.

12 Chile saltpetre, $NaNO_3$, contains sodium iodide as an impurity.

Aqueous silver nitrate is added to an aqueous solution of Chile saltpetre. Concentrated aqueous ammonia is then added.

Which observations are made?

	with acidified silver nitrate	with concentrated aqueous ammonia
A	no precipitate	no further reaction
B	no precipitate	precipitate forms
C	precipitate forms	precipitate dissolves
D	precipitate forms	precipitate remains

13 What happens when chlorine is bubbled through aqueous potassium iodide?

A Chlorine is oxidised to chlorate(V) ions.

B Chlorine is oxidised to chloride ions.

C Iodide ions are oxidised to iodine.

D There is no observable reaction.

14 Which statement about bromine is correct?

A Bromine is insoluble in non-polar solvents.

B Bromine vapour is more dense than air.

C Bromine will not vapourise significantly under normal conditions.

D Gaseous bromine is purple.

15 The following report appeared in a newspaper.

> Drums of bromine broke open after a vehicle crash on the motorway. Traffic was diverted as purple gaseous bromine drifted over the road (it is denser than air), causing irritation to drivers' eyes. Firemen sprayed water over the scene of the accident, dissolving the bromine and washing it away.

What is **wrong** with the report?

A Bromine does not dissolve in water.

B Bromine does not vapourise readily.

C Bromine is less dense than air.

D Bromine is not purple.

17 In what order does the reducing power of the hydrogen halides increase?

A HC*l*, HBr, HI

B HC*l*, HI, HBr

C HBr, HI, HC*l*

D HI, HBr, HC*l*

18 Mixing aqueous silver nitrate and aqueous sodium chloride produces a precipitate.

Addition of which reagent to the mixture gives a colourless solution?

A aqueous ammonia

B aqueous potassium iodide

C dilute hydrochloric acid

D dilute nitric acid

19 In black and white photographic film, light converts silver chloride into metallic silver. After the film has been developed, the unexposed silver chloride is removed by reaction with sodium thiosulphate to produce a 'fixed' negative.

$$AgCl + 2Na_2S_2O_3 \longrightarrow 4Na^+ + Cl^- + [Ag(S_2O_3)_2]^{3-}$$

What is the function of thiosulphate?

A to make the silver ions soluble

B to oxidise the silver ions

C to oxidise the silver metal

D to reduce silver ions

20 X, Y and Z represent different halogens. The table shows the results of nine experiments in which aqueous solutions of X_2, Y_2 and Z_2 were separately added to separate aqueous solutions containing X^-, Y^- and Z^- ions.

	X^-(aq)	Y^-(aq)	Z^-(aq)
X_2(aq)	no reaction	no reaction	no reaction
Y_2(aq)	X_2 formed	no reaction	Z_2 formed
Z_2(aq)	X_2 formed	no reaction	no reaction

Which row in the following table contains the ions X^-, Y^- and Z^- in order of their decreasing strength as reducing agents?

	strongest	$\longrightarrow$	weakest
A	X^-	Y^-	Z^-
B	X^-	Z^-	Y^-
C	Y^-	Z^-	X^-
D	Z^-	X^-	Y^-

21 Which gaseous hydride most readily decomposes into its elements on contact with a hot glass rod?

A ammonia

B hydrogen chloride

C hydrogen iodide

D steam

22 How does ammonia behave when its aqueous solution is used to dissolve silver chloride, $AgCl$?

A as a base

B as a ligand with the Ag^+ ion

C as a ligand with the Cl^- ion

D as a reducing agent

23 Concentrated sulphuric acid is added to separate solid samples of sodium chloride, sodium bromide or sodium iodide.

With which sample(s) does sulphuric acid act as an oxidising agent?

A sodium chloride only

B sodium chloride and sodium bromide

C sodium bromide and sodium iodide

D sodium iodide only

24 A student observed the reactions when sodium chloride and sodium iodide were each reacted separately with concentrated sulfuric acid and with concentrated phosphoric acid. The observations are recorded in the table.

	sodium chloride	sodium iodide
conc. H_2SO_4	colourless acidic gas formed	purple vapour formed
conc. H_3PO_4	colourless acidic gas formed	colourless acidic gas formed

Which deduction can be made from these observations?

A Concentrated phosphoric acid is a stronger oxidising agent than concentrated sulfuric acid.

B Concentrated phosphoric acid is a stronger oxidising agent than iodine.

C Concentrated sulfuric acid is a stronger oxidising agent than chlorine.

D Concentrated sulfuric acid is a stronger oxidising agent than iodine.

25 Concentrated sulfuric acid reacts with both solid sodium chloride at room temperature and with solid sodium iodide at room temperature.

Which row correctly describes how concentrated sulfuric acid behaves in each of these reactions?

	with sodium chloride	with sodium iodide
A	as an oxidising agent only	as an oxidising agent only
B	as a strong acid and as an oxidising agent	as a strong acid only
C	as a strong acid only	as a strong acid and as an oxidising agent
D	as a strong acid only	as a strong acid only

26 Which is the complete list of all the products from the reaction of concentrated sulphuric acid with potassium bromide?

A potassium hydrogensulphate and hydrogen bromide

B potassium hydrogensulphate, hydrogen bromide and bromine

C potassium hydrogensulphate, hydrogen bromide, bromine and water

D potassium hydrogensulphate, hydrogen bromide, bromine, water and sulphur dioxide

27 Chlorine dioxide is produced on a large scale as it is used for bleaching paper pulp. It is made by the following reaction.

$$2ClO_3^-(aq) + SO_2(g) \rightarrow 2ClO_2(g) + SO_4^{2-}(aq)$$

How do the oxidation numbers of chlorine and sulphur change in this reaction?

	chlorine	sulphur
A	decreases by 1	increases by 1
B	decreases by 1	increases by 2
C	decreases by 3	increases by 1
D	decreases by 3	increases by 2

29 In the treatment of domestic water supplies, chlorine is added to the water to form chloric(I) acid, HC*l*O.

$$Cl_2(aq) + H_2O(l) \rightarrow H^+(aq) + Cl^-(aq) + HClO(aq)$$

This reacts further to give the chlorate(I) ion.

$$HClO(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO^-(aq)$$

Both HC*l*O and ClO^- kill bacteria by oxidation.

What is the change in oxidation number of chlorine in forming the chlorate(I) ion from the aqueous chlorine?

A –1 **B** 0 **C** +1 **D** +2

30 A weedkiller can be prepared by heating a bleach solution.

$$\underset{\text{bleach}}{3NaClO} \xrightarrow{\text{heat}} 2NaCl + \underset{\text{weedkiller}}{NaClO_3}$$

What are the oxidation states of chlorine in these three compounds?

A –1 –1 +5

B +1 –1 +5

C +1 –1 +7

D +2 +1 +7

31 In black and white photographic film, light converts silver chloride into metallic silver. After the film has been developed, the unreacted silver chloride is removed by reaction with sodium thiosulfate to produce a ‘fixed’ negative.

$$AgCl + 2Na_2S_2O_3 \rightarrow 4Na^+ + Cl^- + [Ag(S_2O_3)_2]^{3-}$$

What is the function of the thiosulfate ion?

A to make the silver ions soluble

B to oxidise the silver ions

C to oxidise the silver metal

D to reduce the silver ions

32 A student observed the reactions when sodium chloride and sodium iodide were each reacted separately with concentrated sulphuric acid and concentrated phosphoric acid. The observations are recorded in the table.

	sodium chloride	sodium iodide
conc. H_2SO_4	colourless acidic gas formed	purple vapour formed
conc. H_3PO_4	colourless acidic gas formed	colourless acidic gas formed

Which deduction can be made from these observations?

A Concentrated phosphoric acid is a stronger oxidising agent than concentrated sulphuric acid.

B Concentrated phosphoric acid is a stronger oxidising agent than iodine.

C Concentrated sulphuric acid is a stronger oxidising agent than chlorine.

D Concentrated sulphuric acid is a stronger oxidising agent than iodine.

33 Compound **X** on refluxing with aqueous sodium hydroxide gave mixture **Y** which on distillation with acidified potassium dichromate(VI) produced propanone. Mixing **Y** with dilute nitric acid and aqueous silver nitrate gave a cream precipitate.

What could compound **X** be?

A $CH_3CHBrCH_3$

B CH_3CHICH_3

C $CH_3CH_2CH_2Br$

D $CH_3CH_2CH_2I$

34 When chlorine and aqueous sodium hydroxide are heated together the following overall reaction occurs.

$$3Cl_2(aq) + 6NaOH(aq) \rightarrow 5NaCl(aq) + NaClO_3(aq) + 3H_2O(l)$$

What are the oxidation numbers for chlorine in each of the following species?

	Cl_2	$NaCl$	$NaClO_3$
A	0	+1	–5
B	+2	–1	+3
C	0	–1	+5
D	–2	+1	–3

36 Why do the halogens become less volatile as Group VII is descended?

A The halogen-halogen bond energy decreases.

B The halogen-halogen bond length increases.

C The number of electrons in each molecule increases.

D The van der Waals' forces between molecules become weaker.

37 In the treatment of domestic water supplies, chlorine is added to the water to form $HClO$.

$$Cl_2(aq) + H_2O(l) \rightarrow H^+(aq) + Cl^-(aq) + HClO(aq)$$

The $HClO$ reacts further to give ClO^- ions.

$$HClO(aq) + H_2O(l) \rightarrow H_3O^+(aq) + ClO^-(aq)$$

Both $HClO$ and ClO^- kill bacteria by oxidation.

What is the overall change in oxidation number of chlorine when forming the ClO^- ion from the aqueous chlorine?

A –1 **B** 0 **C** +1 **D** +2

38 What trend is observed on descending Group VII?

A The colours of the elements become lighter.

B The elements become more volatile.

C The hydrides of the elements become more thermally stable.

D The reactions of the elements with hydrogen become less vigorous.

39 The following two experiments are carried out with anhydrous potassium chloride and observations X and Y are made at the end of each experiment.

Concentrated sulfuric acid is added to the potassium chloride and the fumes produced are bubbled into aqueous potassium iodide solution - observation X.

The potassium chloride is dissolved in aqueous ammonia and this is then added to aqueous silver nitrate - observation Y.

What are the observations X and Y?

	X	Y
A	brown solution	colourless solution
B	brown solution	white precipitate
C	colourless solution	colourless solution
D	colourless solution	white precipitate

40 Chlorine shows oxidation states ranging from –1 to +7 in its compounds.

What are the reagent(s) and conditions necessary for the oxidation of elemental chlorine into a compound containing chlorine in the +5 oxidation state?

A $AgNO_3$(aq) followed by NH_3(aq) at room temperature

B concentrated H_2SO_4 at room temperature

C cold dilute NaOH(aq)

D hot concentrated NaOH(aq)

41 The reaction between KI and concentrated H_2SO_4 is a redox reaction.

$$5H_2SO_4 + 8KI \rightarrow 4K_2SO_4 + 4I_2 + H_2S + 4H_2O$$

What is the change in oxidation state of the element that is reduced?

A 1 **B** 4 **C** 6 **D** 8

42 Solutions containing chlorate(I) ions are used as household bleaches and disinfectants. These solutions decompose on heating as shown.

$$3ClO^- \rightarrow ClO_3^- + 2Cl^-$$

Which oxidation state is shown by chlorine in each of these three ions?

	ClO^-	ClO_3^-	Cl^-
A	+1	+3	–1
B	–1	+3	+1
C	+1	+5	–1
D	–1	+5	+1

43 What happens when iodine solution is added to a solution of sodium bromide?

A A reaction occurs without changes in oxidation state.

B Bromide ions are oxidised, iodine atoms are reduced.

C Bromide ions are reduced, iodine atoms are oxidised.

D No reaction occurs.

44 Element 85, astatine, is in Group VII. Concentrated sulfuric acid is added to sodium astatide. The mixture of products includes astatine, hydrogen astatide, hydrogen sulfide, and sodium sulfate.

Which product is formed by the oxidation of one of the constituents of sodium astatide?

A astatine

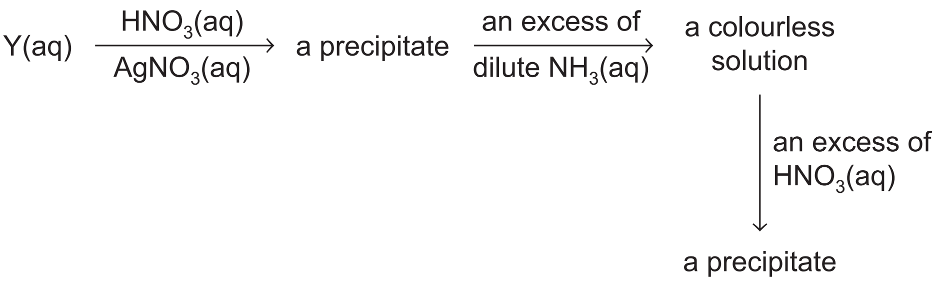
B hydrogen astatide

C hydrogen sulfide

D sodium sulfate

45 Y is a salt of one of the halogens chlorine, bromine, iodine, or astatine (element 85).

The reaction scheme shows a series of reactions using a solution of Y as the starting reagent.



What could Y be?

A sodium chloride

B sodium bromide

C potassium iodide

D potassium astatide

46 Many modern cars are fitted with halogen lamps. When such lamps are first switched on, a distinct purple colour can be seen.

Which species is responsible for this purple colour?

A I_2(s) **B** I_2(l) **C** I_2(g) **D** I(g)

47 What happens when bromine solution is added to a solution of sodium iodide?

A A reaction occurs without changes in oxidation state.

B Bromine atoms are oxidised, iodide ions are reduced.

C Bromine atoms are reduced, iodide ions are oxidised.

D No reaction occurs.

48 The halogens exist as diatomic molecules, X_2.

The boiling points of the Group VII elements increase as the group is descended from chlorine to iodine.

Which statement helps to explain this increase in boiling point as Group VII is descended?

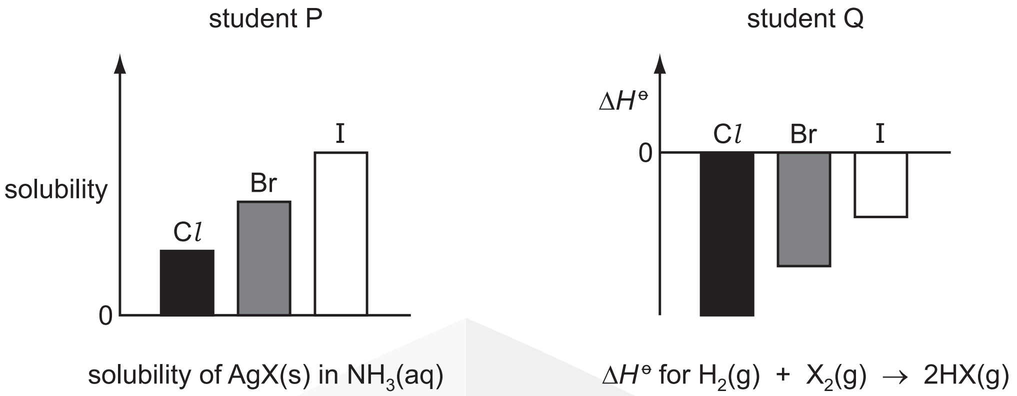
A The electronegativity of X decreases as the group is descended.

B The number of electrons in each X_2 molecule increases as the group is descended.

C The size of the permanent dipole in the X_2 molecule increases as the group is descended.

D The X–X bond strength increases as the group is descended.

49 Two students, P and Q, were asked to draw bar charts to represent how some properties of the halogens and their compounds differ in magnitude. Their diagrams are shown.



Which of the student's diagrams are correct?

A both P and Q

B P only

C Q only

D neither P nor Q

50 When iodine is heated, a vapour is produced.

Which row of the table correctly identifies the species in the vapour and its colour?

	species	colour
A	$I(g)$	brown
B	$I(g)$	purple
C	$I_2(g)$	brown
D	$I_2(g)$	purple

51 How do the strengths of the covalent bonds within molecules, and the van der Waals' forces between molecules, vary going down Group VII from chlorine to bromine to iodine?

	strength of covalent bonds	strength of van der Waals' forces
A	decrease	decrease
B	decrease	increase
C	increase	decrease
D	increase	increase

52 Element X forms X^- ions that can be oxidised to element X by acidified potassium manganate(VII).

What could be the values of the first four ionisation energies of X?

	1st	2nd	3rd	4th
A	418	3070	4600	5860
B	577	1820	2740	11 600
C	590	1150	4940	6480
D	1010	1840	2040	4030

53 Redox reactions occur very frequently in the chemistry of Group VII.

Which statement is correct?

A Chlorine will oxidise bromide ions but not iodide ions.

B Fluorine is the weakest oxidising agent out of F_2, Cl_2, Br_2 and I_2.

C Iodide ions are the weakest reducing agent out of F^-, Cl^-, Br^- and I^-.

D When chlorine reacts with water, chlorine is both oxidised and reduced.

54 When solid sodium iodide reacts with concentrated sulfuric acid, the products include $NaHSO_4$, H_2S, SO_2 and S.

In the formation of which product has the oxidation state of sulfur changed by a value of 8?

A H_2S **B** $NaHSO_4$ **C** S **D** SO_2

55 A test-tube of HBr(g) and a test-tube of HI(g) are heated to the same temperature.

Which combination of observations is possible?

	tube of HBr(g)	tube of HI(g)
A	a brown vapour appears	no change is apparent
B	a purple vapour appears	no change is apparent
C	no change is apparent	a brown vapour appears
D	no change is apparent	a purple vapour appears

56 In an experiment, 0.125 mol of chlorine gas, Cl_2, is reacted with an excess of cold aqueous sodium hydroxide. One of the products is a compound of sodium, oxygen, and chlorine.

Which mass of this product is formed?

A 9.31 g **B** 13.3 g **C** 18.6 g **D** 26.6 g

57 Chlorine gas reacts with cold aqueous sodium hydroxide. It can also react with hot aqueous sodium hydroxide.

What are the oxidation numbers of chlorine in the products of these reactions?

	cold aqueous sodium hydroxide	hot aqueous sodium hydroxide
A	–1, +1	–1, +5
B	–1, +1	+1, +6
C	–1, +2	–1, +5
D	–1, +2	+1, +6

58 Under standard conditions, which statement is correct?

A Cl^-(aq) can oxidise Br_2(aq).

B Cl^-(aq) can reduce Br_2(aq).

C Cl_2(aq) can oxidise Br^-(aq).

D Cl_2(aq) can reduce Br^-(aq).

59 A powder is known to be either a single sodium halide or a mixture of two sodium halides. A sample of the powder was dissolved in water. Aqueous silver nitrate was added, and a pale yellow precipitate was formed. When concentrated aqueous ammonia was added, the precipitate partly dissolved leaving a darker yellow precipitate.

What might the powder have consisted of?

A sodium bromide only

B sodium iodide only

C a mixture of sodium chloride and sodium bromide

D a mixture of sodium chloride and sodium iodide

60 Chlorine is widely used in water treatment plants.

Which reaction takes place when chlorine dissolves in water?

A $Cl_2 + H_2O \rightarrow HCl + HClO$

B $2Cl_2 + 2H_2O \rightarrow 3HCl + HClO_2$

C $3Cl_2 + 3H_2O \rightarrow 5HCl + HClO_3$

D $4Cl_2 + 4H_2O \rightarrow 7HCl + HClO_4$

62 Chlorate(V) ions, ClO_3^-, are produced in the redox reaction between chlorine and hot aqueous sodium hydroxide. Oxidation numbers can be used to help balance the equation for this reaction.

What will be the values of coefficients **v**, **x** and **y** in the balanced equation?

$$\mathbf{v}Cl_2(g) + \mathbf{w}OH^-(aq) \rightarrow \mathbf{x}Cl^-(aq) + \mathbf{y}ClO_3^-(aq) + \mathbf{z}H_2O(l)$$

	v	**x**	**y**
A	2	3	1
B	3	4	2
C	3	5	1
D	7	12	2

63 Two separate tests were carried out on an aqueous solution of a salt which contains only two ions.

In each test, the resulting mixtures were filtered and the filtrate solution was collected.

	reagent	result	filtrate collected
test 1	aqueous silver nitrate acidified with dilute nitric acid	yellow precipitate forms	solution P
test 2	dilute sulfuric acid	white precipitate forms	solution Q

Which ions are present in the filtrates?

	solution P	solution Q
A	$Ba^{2+}(aq)$	$Cl^-(aq)$
B	$Ba^{2+}(aq)$	$I^-(aq)$
C	$Mg^{2+}(aq)$	$Cl^-(aq)$
D	$Mg^{2+}(aq)$	$I^-(aq)$

64 Which statement about bromine is correct?

A Bromine is insoluble in non-polar solvents.

B Bromine vapour is more dense than air.

C Bromine will not vaporise significantly under normal conditions.

D Gaseous bromine is purple.

65 The addition of aqueous silver nitrate to aqueous barium chloride produces a white precipitate which dissolves in excess dilute aqueous ammonia to form a colourless solution.

The addition of excess dilute nitric acid to the colourless solution produces a white precipitate, **Z**.

What is **Z**?

A $AgCl$ **B** $BaCl_2$ **C** $Ba(NO_3)_2$ **D** NH_4NO_3

66 A student observed the reactions when sodium chloride and sodium iodide were each reacted separately with concentrated sulfuric acid and with concentrated phosphoric acid. Some observations are recorded in the table.

	sodium chloride	sodium iodide
conc. H_2SO_4	colourless acidic gas formed	purple vapour formed
conc. H_3PO_4	colourless acidic gas formed	colourless acidic gas formed

Which deduction can be made from these observations?

A Concentrated phosphoric acid is a stronger oxidising agent than concentrated sulfuric acid.

B Concentrated phosphoric acid is a stronger oxidising agent than iodine.

C Concentrated sulfuric acid is a stronger oxidising agent than chlorine.

D Concentrated sulfuric acid is a stronger oxidising agent than iodine.

67 A white powder is a mixture of sodium chloride and sodium iodide. It is dissolved in water in a test-tube. Excess aqueous silver nitrate is added to the test-tube. A precipitate, X, is observed.

Excess concentrated ammonia is then added to the test-tube containing X. After the test-tube has been shaken, a precipitate, Y, is observed.

Which statement about X or Y is correct?

A X is a pure white colour.

B X is pure silver iodide.

C Y is pure silver chloride.

D Y is yellow.

68 Chlorine gas is widely used to treat contaminated water.

Which species present in water when chlorine gas has been added is responsible for killing bacteria?

A ClO_2^- **B** Cl^- **C** HCl **D** OCl^-

69 Hydrogen chloride gas and hydrogen iodide gas have different thermal stabilities. The difference is due to a difference in the energies of some of the covalent bonds that are involved in the decomposition.

Which row identifies the more stable of the two compounds, and gives the correct explanation?

	identity of the more thermally stable compound	explanation for the difference in stability
A	hydrogen chloride	the Cl–Cl bond is stronger than the I–I bond
B	hydrogen chloride	the H–Cl bond is stronger than the H–I bond
C	hydrogen iodide	the Cl–Cl bond is stronger than the I–I bond
D	hydrogen iodide	the H–Cl bond is stronger than the H–I bond

70 The solids sodium chloride and sodium iodide both react with concentrated sulfuric acid at room temperature.

With $NaCl$, the products are $NaHSO_4$ and HCl.
With NaI, the products are $NaHSO_4$, HI, I_2, SO_2, H_2O, S and H_2S.

What is the best explanation for this difference in products?

A Chloride ions will displace iodine from solution.

B Hydrogen chloride is more volatile than hydrogen iodide.

C Iodide ions are better reducing agents than chloride ions.

D Sulfuric acid is able to act as a dehydrating agent with NaI.

71 In the treatment of domestic water supplies, chlorine is added to the water to form $HClO$.

$$Cl_2(aq) + H_2O(l) \rightarrow H^+(aq) + Cl^-(aq) + HClO(aq)$$

The $HClO$ reacts further to give ClO^- ions.

$$HClO(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + ClO^-(aq)$$

Both $HClO$ and ClO^- kill bacteria by oxidation.

What is the change in oxidation number of chlorine when forming the ClO^- ion from aqueous chlorine?

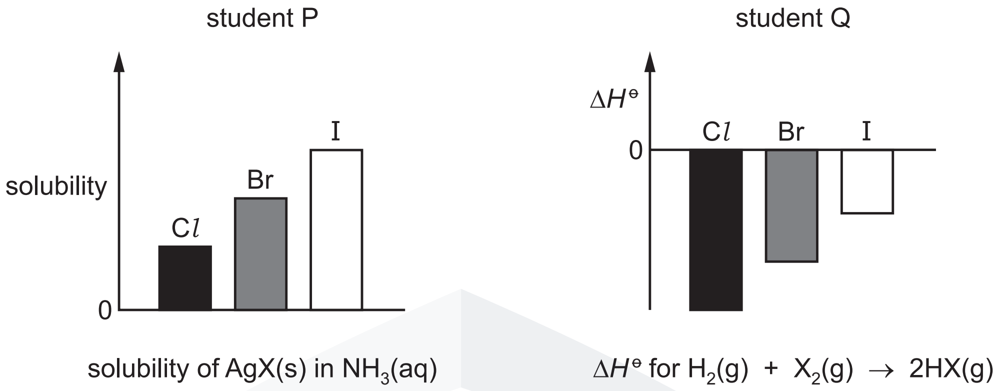
A –1 **B** 0 **C** +1 **D** +2

72 Fluorine and iodine are Group 17 elements. Their melting points are different due to differing strengths of van der Waals' forces between molecules.

Which row is correct?

	melting point	strength of van der Waals' forces between molecules
A	$F_2 > I_2$	$F_2 > I_2$
B	$F_2 > I_2$	$F_2 < I_2$
C	$F_2 < I_2$	$F_2 < I_2$
D	$F_2 < I_2$	$F_2 > I_2$

74 Two students, P and Q, were asked to draw bar charts to represent how some properties of the halogens and their compounds differ in magnitude. Their diagrams are shown. The bar charts show trends only and not actual values.



Which of the students have drawn bar charts which show the trends correctly?

A both P and Q

B P only

C Q only

D neither P nor Q

75 The properties of chlorine, bromine and their compounds are compared.

Which property is **smaller** for chlorine than for bromine?

A bond strength of the hydrogen-halide bond

B first ionisation energy

C solubility of the silver halide in NH_3(aq)

D strength of the van der Waals' forces between molecules of the element

77 PVC is difficult to dispose of. Two possible methods are burying it in landfill sites and disposal by combustion.

Which row of the table is correct?

	rate of biodegradation of PVC in landfill sites	gases produced when PVC combusts
A	fast	CO_2, H_2O, HCl
B	fast	CO_2, H_2O, Cl_2
C	slow	CO_2, H_2O, Cl_2
D	slow	CO_2, H_2O, HCl

78 Chlorine and bromine have different volatilities.

Which row identifies the more volatile of the two elements, and gives the correct explanation?

	identity of the more volatile element	explanation for the difference in volatility
A	bromine	instantaneous dipole-induced dipole forces are greater in bromine than they are in chlorine
B	bromine	instantaneous dipole-induced dipole forces are greater in chlorine than they are in bromine
C	chlorine	instantaneous dipole-induced dipole forces are greater in bromine than they are in chlorine
D	chlorine	instantaneous dipole-induced dipole forces are greater in chlorine than they are in bromine

79 The properties of chlorine, bromine and their compounds are compared.

Which property is **smaller** for chlorine than for bromine?

A bond strength of the hydrogen-halide bond

B first ionisation energy

C solubility of the silver halide in $NH_3(aq)$

D strength of the van der Waals' forces between molecules of the element

80 A solution of sodium hydroxide reacts with 3 mol of chlorine under certain conditions. The reaction produces 5 mol of sodium chloride and 1 mol of **X**, the only other chlorine-containing product.

What is the formula of compound **X**?

A $NaClO$ **B** $NaClO_2$ **C** $NaClO_3$ **D** $NaClO_4$

81 A test-tube of HBr(g) and a separate test-tube of HI(g) are heated to the same temperature.

Which combination of observations is possible?

	test-tube of HBr(g)	test-tube of HI(g)
A	a brown vapour appears	no change
B	a purple vapour appears	no change
C	no change	a brown vapour appears
D	no change	a purple vapour appears

82 Some uses of chlorine and bromine are given.

Which is a use of bromine?

A making bleaches for textiles and the paper industry

B making CFCs

C making flame retardants and fire extinguishers

D making the polymer PVC

83 In an experiment, 0.125 mol of chlorine gas, Cl_2, is reacted with an excess of cold, aqueous sodium hydroxide. One of the products is a compound of sodium, oxygen and chlorine.

Which mass of this product is formed?

A 9.31 g **B** 13.3 g **C** 18.6 g **D** 26.6 g

[W'17 2 Q16]

84 Sodium bromide reacts with concentrated sulfuric acid.

Which observation will be made?

A A coloured vapour is produced.

B A purple solid is formed.

C A strong smell of H_2S is detected.

D Yellow sulfur is formed.

[W'17 2 Q17]

85 71.0 g of chlorine, Cl_2, react with an excess of sodium hydroxide solution at a particular temperature. The reaction produces exactly 35.5 g of product **X**.

What is product **X**?

A H_2O **B** $NaCl$ **C** $NaClO$ **D** $NaClO_3$

[W'17 3 Q17]

87 Bromine is extracted from sea-water.

In the final stages of the process two redox reactions take place.

$$Br_2(aq) + SO_2(g) + 2H_2O(l) \rightarrow 2HBr(aq) + H_2SO_4(aq)$$

$$2HBr(aq) + Cl_2(g) \rightarrow Br_2(g) + 2HCl(aq)$$

Which row is correct?

	strongest oxidising agent	→	weakest oxidising agent
A	Br_2	SO_2	Cl_2
B	Cl_2	Br_2	SO_2
C	Cl_2	SO_2	Br_2
D	SO_2	Br_2	Cl_2

[M'18 2 Q17]

88 Chlorine reacts with **cold** aqueous sodium hydroxide to produce sodium chloride, water and compound X.

Chlorine reacts with **hot** aqueous sodium hydroxide to produce sodium chloride, water and compound Y.

What are the oxidation states of chlorine in compound X and compound Y?

	X	Y
A	–1	–5
B	–1	+5
C	+1	–5
D	+1	+5

[S'18 1 Q14]

89 Aqueous silver nitrate is added to a solution of potassium iodide.

Aqueous ammonia is then added.

What would be observed?

A a cream precipitate that dissolves on addition of aqueous ammonia

B a cream precipitate that does not dissolve on addition of aqueous ammonia

C a yellow precipitate that dissolves on addition of aqueous ammonia

D a yellow precipitate that does not dissolve on addition of aqueous ammonia

[S'18 1 Q16]

90 When concentrated sulfuric acid is added to solid sodium bromide, bromine gas is produced, along with a number of other products. However when concentrated sulfuric acid is added to solid sodium chloride **only** hydrogen chloride and sodium hydrogensulfate are produced.

What is the reason for this difference?

A Bromine is less volatile than chlorine.

B Hydrochloric acid is a weak acid.

C Sulfuric acid is not an oxidising agent.

D The bromide ion is a stronger reducing agent than the chloride ion.

[S'18 2 Q18]

SECTION B

For the following questions, the responses A to D should be selected on the basis of

A	B	C	D
1, 2 and **3** are correct	**1** and **2** only are correct	**2** and **3** only are correct	**1** only is correct

1 Which sodium salts form a precipitate when $AgNO_3$(aq) followed by dilute NH_3(aq) is added to its aqueous solution?

1 chloride

2 bromide

3 iodide

2 Chlorine is a greenish-yellow gas, bromine is a dark red liquid and iodine is a dark grey solid.

What causes these differences in volatility?

1 the halogen-halogen bond energy

2 the magnitude of the van der Waals' forces between the molecules

3 the number of electrons in the halogen molecule

3 The element astatine lies below iodine in Group VII of the Periodic Table.

What will be the properties of astatine?

1 It forms diatomic molecules which dissociate more readily than chlorine molecules.

2 It reacts explosively with hydrogen.

3 It is a good reducing agent.

4 What happens when chlorine is bubbled through aqueous sodium hydroxide solution?

1 In cold solution, ClO^-(aq) ions are formed.

2 In hot solution, ClO_3^-(aq) ions are formed.

3 Disproportionation of chlorine occurs in both cold and hot aqueous solutions.

5 The element astatine lies below iodine in Group VII of the Periodic Table.

What will be the properties of astatine?

1 It forms diatomic molecules which dissociate more readily than chlorine molecules.

2 It reacts explosively with hydrogen.

3 It is a good reducing agent.

6 Why is the addition of concentrated sulphuric acid to solid potassium iodide **unsuitable** for the preparation of hydrogen iodide?

1 Hydrogen iodide is not displaced by sulphuric acid.

2 Iodide ions are oxidised to iodine.

3 The product is contaminated by sulphur compounds.

7 When a hot glass rod is placed in a gas jar of hydrogen iodide, there is an immediate reaction as the hydrogen iodide decomposes.

Which statements about this reaction are correct?

1 Hydrogen iodide is purple coloured.

2 The hot rod provides the activation energy.

3 One of the products is a solid.

8 Which properties in the sequence hydrogen chloride, hydrogen bromide and hydrogen iodide steadily increase?

1 thermal stability

2 bond length

3 ease of oxidation

9 The element astatine lies below iodine in Group VII of the Periodic Table.

What will be the properties of astatine?

1 It forms diatomic molecules which dissociate more readily than chlorine molecules.

2 It reacts explosively with hydrogen.

3 It can oxidise iodide to iodine.

10 Which statements about the reaction of solid sodium bromide with concentrated sulfuric acid are correct?

1 Hydrogen bromide is a product of the reaction.

2 Sulfuric acid is oxidised to sulfur dioxide.

3 Bromide ions are reduced to bromine.

11 Disproportionation is the term used to describe a reaction in which a reactant is simultaneously both oxidised and reduced.

To which incomplete equations does the term disproportionation apply?

1 $Cl_2(g) + 2OH^-(aq) \rightarrow H_2O(l) + Cl^-(aq) + \ldots\ldots$

2 $3Cl_2(g) + 6OH^-(aq) \rightarrow 3H_2O(l) + ClO_3^-(aq) + \ldots\ldots$

3 $2NO_2(g) + H_2O(l) \rightarrow HNO_3(aq) + \ldots\ldots$

12 The number of moles of chlorine that react with 1 mol of *X* is twice the number of moles of chlorine that react with 1 mol of *Y*.

Which of these pairs could be *X* and *Y*?

	X	*Y*
1	Mg(s)	Na(s)
2	H_2	KBr(aq)
3	cold NaOH(aq)	hot NaOH(aq)

13 When the yellow liquid NCl_3 is stirred into aqueous sodium hydroxide, the reaction that occurs can be represented by the following equation.

$$2NCl_3(l) + 6NaOH(aq) \rightarrow N_2(g) + 3NaCl(aq) + 3NaOCl(aq) + 3H_2O(l)$$

What will be the result of this reaction?

1 The nitrogen is oxidised.

2 A bleaching solution remains after the reaction.

3 The final solution gives a precipitate with acidified silver nitrate.

14 Why is the addition of concentrated sulphuric acid to solid potassium iodide **unsuitable** for the preparation of hydrogen iodide?

1 Hydrogen iodide is not displaced by sulphuric acid.

2 Iodide ions are oxidised to iodine.

3 The product is contaminated by sulphur compounds.

15 Which of the halide ions, chloride, bromide or iodide, acts as a reducing agent when its sodium salt reacts with concentrated sulfuric acid?

1 at least one of Cl^-, Br^- and I^-

2 at least two of Cl^-, Br^- and I^-

3 all three of Cl^-, Br^- and I^-

16 Which statements are correct for all three halogens, chlorine, bromine and iodine?

1 They all form hydrides that are strong acids in aqueous solution.

2 They all react with aqueous sodium hydroxide to form oxo-anions.

3 They all require one more electron to fill the p orbitals of their outer shells.

17 Compared with the HCl molecule, the bond ……**X**…… of the HBr molecule is ……**Y**……

Which pairs of words correctly complete the above sentence?

	X	**Y**
1	energy	less
2	polarity	less
3	length	greater

18 The element astatine, At, is below iodine in Group VII of the Periodic Table.

Which statements concerning At will be correct?

1 It is a dark-coloured solid at room temperature.

2 It is a more powerful oxidising agent than iodine.

3 Its hydride is thermally stable.

19 The table describes some of the chemistry and thermodynamic properties of the halogens.

process	name and symbol of quantity
$2HX(g) \rightarrow H_2(g) + X_2(g)$	enthalpy change of reaction, $\Delta H^{\ominus}$
$H_2(g) + X_2(g) \rightleftharpoons 2HX(g)$	equilibrium constant, K_p
$X(g) \rightarrow X^+(g) + e^-$	ionisation energy, $\Delta H_i^{\ominus}$

Which statements about the relative values of these quantities are correct?

1 $\Delta H^{\ominus}$ for HCl > $\Delta H^{\ominus}$ for HBr

2 K_p for HBr > K_p for HI

3 $\Delta H_i^{\ominus}$ for I > $\Delta H_i^{\ominus}$ for Cl

20 When a red-hot platinum wire is plunged into a test tube of hydrogen iodide, the gas is decomposed into its elements. If the experiment is repeated with hydrogen chloride, no change occurs.

Which factors contribute to this behaviour?

1 the strength of the hydrogen-halogen bond

2 the size of the halogen atom

3 the standard enthalpy of formation, $\Delta H_f^{\ominus}$, of each of the products of decomposition

21 Which properties increase in the sequence hydrogen chloride, hydrogen bromide and hydrogen iodide?

1 thermal stability

2 bond length

3 ease of oxidation

22 A test-tube of HI gas and a test-tube of HBr gas are placed together in an environment at temperature, T.

Which combinations of observations are possible depending on the temperature, T?

1 A brown vapour appears in one of the test-tubes. No change is apparent in the other test-tube.

2 A brown vapour appears in one of the test-tubes. A purple vapour appears in the other test-tube.

3 No change is apparent in either test-tube.

23 Three samples of chlorine gas each contain 0.710 g of chlorine. Each sample is reacted with a reagent.

- In the first reaction a sample is reacted completely with hydrogen gas.
- In the second reaction a sample is reacted completely with cold NaOH(aq).
- In the third reaction a sample is reacted completely with hot NaOH(aq).

Which masses of the named products would be formed?

1 Exactly 0.730 g of HCl form in the first reaction.

2 Exactly 0.585 g of $NaCl$ form in the second reaction.

3 Exactly 0.975 g of $NaCl$ form in the third reaction.

[S'18 2 Q35]

24 The intermolecular forces between iodine molecules are instantaneous dipole-induced dipole forces.

Which statements explain why iodine has these intermolecular forces?

1 An iodine molecule is polar and experiences an attraction from a lone pair of electrons on an adjacent molecule.

2 An iodine molecule has a fluctuating dipole because the electrons in a molecule are more mobile than the nuclei.

3 The electron charge cloud within an I_2 molecule may become unsymmetrical and may then attract other I_2 molecules.

25 On being heated, hydrogen iodide breaks down more quickly than hydrogen chloride.

Which statements explain this faster rate?

1 The HI bond is weaker than the HCl bond.

2 The reaction of the breakdown of HI has a smaller activation energy than that of HCl.

3 The breakdown of HI is more exothermic than that of HCl.

26 Each of the three mixtures shown can result in a chemical reaction.

Which mixtures result in a redox reaction?

1 bromine + hydrogen

2 sodium chloride + concentrated sulfuric acid

3 potassium iodide + silver nitrate

[S'18 1 Q35]

27 The salt $NaClO_3$ is used as a non-selective weedkiller.

On careful heating, this reaction occurs: $4NaClO_3 \rightarrow NaCl + 3NaClO_4$.

On strong heating this reaction occurs: $NaClO_4 \rightarrow NaCl + 2O_2$.

The overall reaction is $2NaClO_3 \rightarrow 2NaCl + 3O_2$.

What do these equations show?

1 $NaClO_3$ can behave as an oxidising agent.

2 $NaClO_3$ can behave as a reducing agent.

3 The oxidation numbers of chlorine in the three compounds shown are +6, +8 and –1.

28 The element astatine, At, is below iodine in Group 17 of the Periodic Table.

Which statements concerning At are likely to be correct?

1 It is a dark-coloured solid at room temperature.

2 It is a more powerful oxidising agent than iodine.

3 Its hydride is thermally stable.

29 When $KClO_3$ is heated, the following reaction occurs.

$$4KClO_3 \rightarrow 3KClO_4 + KCl$$

Which statements are correct?

1 The oxidation state of Cl in $KClO_3$ is +5.

2 The oxidation state of some Cl atoms decreases by 6.

3 The reaction involves disproportionation.

30 Chlorine reacts with **hot** aqueous sodium hydroxide.

Which oxidation states does chlorine show in the products of this reaction?

1 –1

2 +3

3 +1

[M'18 2 Q35]

NOTES

GROUP 17 WS 2

1 This question is about the elements of Group VII, the halogens.

(a) Complete the following table.

halogen	colour	physical state at room temperature
chlorine		
bromine		
iodine		

[2]

(b) Concentrated sulphuric acid is added to separate solid samples of magnesium chloride, magnesium bromide, and magnesium iodide.

(i) Describe, in **each** case, **one** observation you would be able to make.

$MgCl_2$ ....

....

$MgBr_2$ ....

....

MgI_2 ....

....

(ii) Give an equation for the reaction of concentrated sulphuric acid with magnesium chloride.

....

[4]

(c) When dilute nitric acid and aqueous silver nitrate are added to a solution of a magnesium halide, MgX_2, a pale cream precipitate is formed.
This precipitate is soluble in concentrated aqueous ammonia but not soluble in dilute aqueous ammonia.

(i) What is the identity of the precipitate?

....

(ii) Give an equation, with state symbols, for the reaction of the precipitate with concentrated aqueous ammonia.

....

[3]

(d) A hot glass rod is plunged into separate gas jars, one containing hydrogen chloride and one containing hydrogen iodide.

(i) For **each** gas, state what you would observe, if anything, and write an equation for any reaction that takes place.

HCl ....

....

HI ....

....

(ii) Explain your answer to **(i)** in terms of enthalpy changes.

....

....

....

(iii) What is the role of the hot glass rod in any reaction that occurs?

....

....

[6]

2 The gaseous hydrogen halides HCl, HBr and HI, may be prepared by reacting the corresponding sodium salt with anhydrous phosphoric(V) acid, H_3PO_4.

When the sodium halide Na*X* was used, the following reaction occurred and a sample of gaseous H*X* was collected in a gas jar.

$$NaX + H_3PO_4 \rightarrow NaH_2PO_4 + HX$$

A hot glass rod was placed in the sample of H*X* and immediately a red/orange colour was observed.

(a) What is the identity of Na*X*?

.... [1]

(b) What gas, other than H*X*, would be formed if concentrated sulfuric acid were used with Na*X* instead of phosphoric(V) acid?

.... [1]

(c) Suggest why phosphoric(V) acid rather than concentrated sulfuric acid is used to make samples of H*X* from the corresponding sodium salt.
Explain your answer.

....

....[1]

3 One method of making 1-bromobutane in the laboratory is described below.

Stage 1	Place 35 g of powdered sodium bromide, 30 cm^3 of water, and 25 cm^3 (20 g) of butan-1-ol, in a 250 cm^3 two necked flask fitted with a tap funnel and reflux condenser.
Stage 2	Concentrated sulfuric acid (25 cm^3) is then placed in the tap funnel and added drop by drop to the reagents in the flask, keeping the contents well shaken and cooled occasionally in an ice-water bath.

(a) The overall reaction may be considered to take place in two stages. In the first stage the inorganic reagents react together to form HBr. In the second stage, the organic reagent reacts with the HBr that is formed in the first stage.

Write an equation for **each** of these stages.

stage I ....

stage II .... [2]

(b) In this preparation, by using the amounts given above, **one** of the reagents, sodium bromide or butan-1-ol, will be present in an excess.

Use your equations in **(a)** and the data above to determine, by calculation, which reagent is in an excess.

[2]

(c) In a laboratory preparation of 1-bromobutane, when 15.4 g of butan-1-ol was used, 22.5 g of 1-bromobutane was obtained after purification.

Calculate the yield of 1-bromobutane as a percentage of the theoretical maximum yield.

[2]

(d) When the concentrated sulfuric acid is added to the reaction mixture (stage 2), unless the temperature is controlled carefully, the acid may react with either of the original reactants (sodium bromide or butan-1-ol) to give at least two by-products, one of which is inorganic and the other organic.

What inorganic and organic by-products may be formed?

In **each** case, identify **one** by-product and state the role of the concentrated sulfuric acid in the formation of this by-product.

inorganic by-product ....

role of conc. H_2SO_4 ....

organic by-product ....

role of conc. H_2SO_4 .... [4]

4 Concentrated sulfuric acid may be used in a school or college laboratory to produce hydrogen chloride by reaction with solid chlorides such as sodium chloride.

(a) **(i)** What will be seen when concentrated sulfuric acid is carefully added to solid sodium chloride?

....

(ii) Write a balanced equation for this reaction.

....

(iii) Solutions of both H_2SO_4 and HCl are strong acids.
What is meant by the term *strong acid*?

....

....

[3]

(b) If the same reaction is carried out with solid sodium iodide and concentrated sulfuric acid, hydrogen iodide is **not** produced.

(i) State **one** observation you would make when carrying out this reaction with solid sodium iodide.

....

(ii) Explain why hydrogen iodide is **not** a product of this reaction.

....

....

[3]

(c) Aqueous silver nitrate and aqueous ammonia are used to test for the presence of halide ions.

(i) Aqueous silver nitrate is slowly added to aqueous sodium chloride and the resulting mixture is then shaken with an excess of aqueous ammonia.

Describe what you would observe at **each** stage of this process.

....

....

....

....

(ii) Write balanced equations, with state symbols, for **all** reactions that occur in this process.

(iii) The same process of adding aqueous silver nitrate followed by an excess of aqueous ammonia is repeated using aqueous sodium iodide instead of aqueous sodium chloride.

State **two** differences that would be observed with aqueous sodium iodide.

....

....

[8]

5 The elements of Group VII of the Periodic Table show variation in their properties.

(a) **(i)** Complete the table below, stating the colour of each element in its normal state at room temperature.

halogen	melting point / °C	colour
chlorine	–101	
bromine	–7	
iodine	114	

(ii) Briefly explain why the melting points of the halogens increase from chlorine to iodine.

....

....

.... [4]

(b) The halogens form many interhalogen compounds in which two different halogens are combined. One such compound is bromine monochloride, BrC*l*.

(i) Complete the electronic configurations of chlorine and bromine.

chlorine	$1s^2 2s^2 2p^6$
bromine	$1s^2 2s^2 2p^6$

(ii) Draw a 'dot-and-cross' diagram of the BrC*l* molecule.
Show outermost electrons only.

[2]

(c) Interhalogen compounds like $BrCl$ have similar properties to the halogens.

(i) By considering your answers to **(a)** and **(b)**, predict the physical state of $BrCl$ at room temperature. Explain your answer.

physical state ....

explanation ....

....

....

(ii) Suggest the colour of $BrCl$.

....

[4]

(d) Cl_2 and $BrCl$ each react with aqueous KI.

(i) Describe what would be seen when Cl_2 is bubbled through aqueous KI for several minutes.

initially ....

....

after several minutes ....

....

(ii) Construct an equation for the reaction that occurs.

....

(iii) Suggest an equation for the reaction that occurs between $BrCl$ and aqueous KI.

....

(iv) How do Cl_2 and $BrCl$ behave in these reactions?

....

[5]

6 **(a)** In this question, **K**, **L** and **M** refer to a halogen atom or halide ion.
For each part question, read the information and complete the answer lines below.

(i) When concentrated sulfuric acid is added to solid Na**K**, white fumes are produced that turn damp blue litmus paper red. No other colour changes are observed.

identity of **K** = ...

equation for reaction ....

explanation of observation ....

.... [3]

(ii) When silver nitrate solution is added to an aqueous solution of Na**L**, a precipitate forms that remains after the addition of concentrated ammonia solution.

identity of **L** = ...

colour of precipitate ....

equation for reaction .... [3]

(iii) $\mathbf{M}_2$ is a liquid at room temperature with a boiling point higher than that of chlorine but lower than that of iodine.

identity of **M** = ...

explanation ....

....

....

....

.... [2]

7 The halogens and their compounds have a wide variety of uses and the chemical and physical properties of the elements show regular patterns related to their positions in Group VII.

(a) Chlorine, bromine and iodine all react with hydrogen.

(i) State the trend in the reactivities of the halogens with hydrogen.

....

.... [1]

(ii) Explain this trend in terms of bond energies.

....

....

.... [2]

(b) In the laboratory it is not very convenient to prepare hydrogen halides from their elements.

Hydrogen halides can be prepared from their salts.

(i) Write an equation for the reaction of calcium chloride, $CaCl_2$, with concentrated sulfuric acid.

.... [1]

(ii) Explain why hydrogen iodide is not prepared in this way.

....

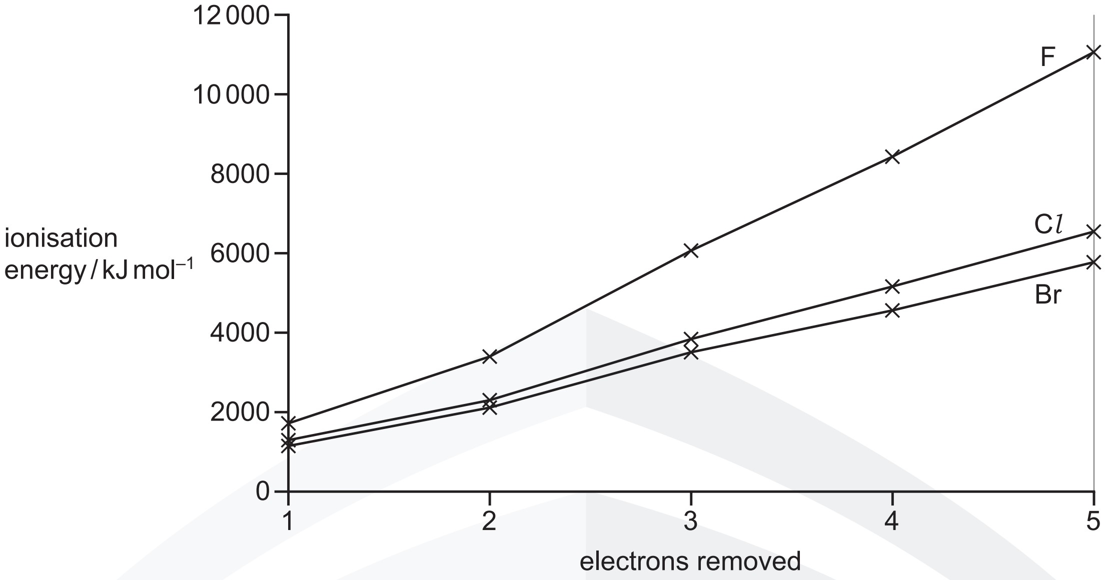
.... [1]

(iii) When potassium bromide, KBr, reacts with concentrated sulfuric acid, sulfur dioxide, SO_2, is produced. State what you would **see** and write an equation for this reaction.

....

.... [3]

8 **(a)** Successive ionisation energies for the elements fluorine, F, to bromine, Br, are shown on the graph.



(i) Explain why the first ionisation energies decrease down the group.

....

....

....

.... [3]

(ii) Explain why there is an increase in the successive ionisation energies of fluorine.

....

....

.... [2]

(b) Group VII is the only group in the Periodic Table containing elements in all three states of matter at room conditions.

State and explain, in terms of intermolecular forces, the trend in the boiling points of the elements down Group VII.

....

....

....

....

....

.... [4]

(c) Compounds containing different halogen atoms covalently bonded together are called interhalogen compounds.

(i) One interhalogen compound can be prepared by the reaction between iodine and fluorine. This compound has M_r = 222 and the percentage composition by mass: F, 42.8; I, 57.2.

Calculate the molecular formula of this interhalogen compound.

molecular formula .... [3]

(ii) Another interhalogen compound has the formula ICl.

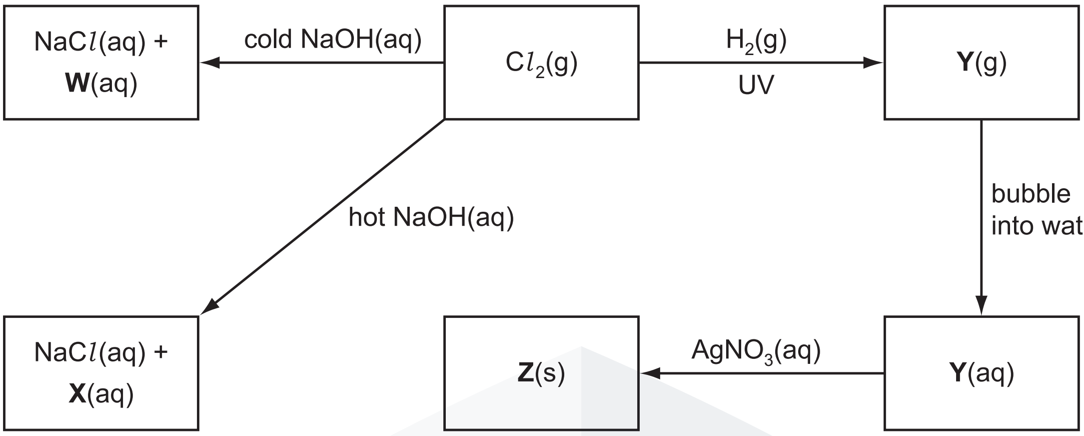
Draw a 'dot-and-cross' diagram of a molecule of this compound, showing outer shell electrons only. Explain whether or not you would expect this molecule to be polar.

....

....

.... [2]

(d) Some reactions involving chlorine and its compounds are shown in the reaction scheme below.



(i) Give the **formulae** of **W**, **X**, **Y** and **Z**.

W ....

X ....

Y ....

Z ....

[4]

(ii) Write an equation for the reaction of chlorine with **hot** NaOH(aq).

.... [2]

(iii) State the oxidation numbers of chlorine at the start and at the end of the reaction in **(ii)**.

.... [2]

(iv) Write an **ionic** equation for the reaction of **Y** with $AgNO_3$(aq). Include state symbols.

.... [1]

9 The elements in Group 17, the halogens, and their compounds, show many similarities and trends in their properties. Some data are given for the elements fluorine to iodine.

element	bond energy / kJ mol^{-1}	standard enthalpy change of atomisation, $\Delta H^{\ominus}_{at}$ / kJ mol^{-1}	boiling point of element / K	boiling point of **hydrogen halide** / K
fluorine, F–F	158	79	85	293
chlorine, Cl–Cl	242	121	238	188
bromine, Br–Br	193	112	332	206
iodine, I–I	151	107	457	238

(a) **(i)** Explain the meaning of the term *standard enthalpy change of atomisation*.

....

....

.... [3]

(ii) For fluorine and chlorine, the enthalpy changes of atomisation are half the value of the bond energies.

For bromine and iodine, the enthalpy changes of atomisation are much more than half the value of the bond energies.

Suggest a reason for this difference.

....

....

.... [1]

(iii) The standard enthalpy of formation of iodine monochloride, ICl, is –24.0 kJ mol^{-1}.

Use this information and the bond energies of iodine and chlorine to calculate the I–Cl bond energy.

I–Cl bond energy = ... kJ mol^{-1} [2]

(b) **(i)** Explain the trend in the boiling points of the hydrogen halides, HC*l*, HBr and HI.

....

....

.... [2]

(ii) Suggest why the hydrogen halide HF does not follow the trend in boiling points shown by HC*l*, HBr and HI.

....

....

.... [2]

(c) In an experiment, two of the halogens are represented as $\mathbf{P}_2$ and $\mathbf{Q}_2$.

$\mathbf{P}_2$ combines with hydrogen on heating to form H**P**, which can be easily broken down into its elements. A solution of H**P** in water reacts with aqueous silver ions to form a yellow precipitate that is insoluble in dilute aqueous ammonia.

$\mathbf{Q}_2$ combines explosively with hydrogen in sunlight to form H**Q**, which is stable to heat. A solution of H**Q** in water reacts with aqueous silver ions to form a white precipitate that is soluble in dilute aqueous ammonia.

(i) Identify the halogens $\mathbf{P}_2$ and $\mathbf{Q}_2$.

$\mathbf{P}_2$ = ... $\mathbf{Q}_2$ = ...
[1]

(ii) H**P** readily decomposes into its elements when heated but H**Q** is stable to heat.
Explain this with reference to bond energies.

....

....

.... [2]

(iii) Write an equation for the thermal decomposition of H**P**.

.... [1]

10 The halogens, chlorine, bromine and iodine, and their compounds, show a variety of similarities and trends in their physical and chemical properties.

(a) **(i)** Give the colours and states of chlorine, bromine and iodine at room temperature and pressure.

halogen	colour	state
chlorine		
bromine		
iodine		

[2]

(ii) The halogens become less volatile down the group.

Explain this trend in volatility.

....

....

.... [2]

(b) The halogens are oxidising agents.

State and explain the trend in oxidising power of the halogens.

....

....

....

.... [3]

(c) Concentrated sulfuric acid reacts with solid sodium halides.

(i) State any observations that would be made on addition of concentrated sulfuric acid to

- solid sodium chloride, ...

....

- solid sodium iodide. ....

....

[2]

(ii) Give reasons for the difference in the observations in **(i)**.

....

....

.... [2]

(iii) The addition of concentrated sulfuric acid to solid sodium bromide, NaBr, produces brown fumes and an acidic gas that decolourises acidified potassium manganate(VII) solution. This acidic gas is a significant contributor to acid rain.

Write the equation for the reaction of concentrated sulfuric acid with sodium bromide.

.... [2]

(d) An aqueous solution, **Z**, contains a mixture of sodium chloride and sodium iodide.

(i) Excess aqueous silver nitrate is added to **Z** in a test-tube. A yellow precipitate forms.

Explain the colour of this precipitate.

....

.... [1]

(ii) Aqueous ammonia is then added to the test-tube in **(i)**. The mass of precipitate decreases.

Explain this observation.

....

.... [1]

11 Hydrogen halides are compounds formed when halogens (Group 17 elements) react with hydrogen. The bond polarity of the hydrogen halides decreases from HF to HI.

Some relevant data are shown in the table.

hydrogen halide	HF	HC*l*	HBr	HI
boiling point / °C	19	–85	–67	–35
H–X bond energy / kJ mol^{-1}	562	431	366	299

(a) (i) Explain the meaning of the term *bond polarity*.

....

....

.... [1]

(ii) Suggest why the boiling point of HF is **much** higher than the boiling points of the other hydrogen halides.

....

....

....

....

.... [2]

(iii) Describe and explain the relative thermal stabilities of the hydrogen halides.

....

....

....

....

....

.... [3]

12 **(b)** The equation for the preparation of hydrogen chloride using concentrated sulfuric acid is shown.

$$H_2SO_4 + NaCl \rightarrow NaHSO_4 + HCl$$

(ii) Explain why the reaction of concentrated sulfuric acid and sodium iodide is **not** suitable for the preparation of hydrogen iodide.

....

....

....

.... [2]

[M'17 2 Q2]

16 The halogens, chlorine, bromine and iodine, and their compounds, show a variety of similarities and trends in their physical and chemical properties.

(a) **(i)** Give the colours and states of chlorine, bromine and iodine at room temperature and pressure.

halogen	colour	state
chlorine		
bromine		
iodine		

[2]

(ii) The halogens become less volatile down the group.

Explain this trend in volatility.

....

....

.... [2]

(b) The halogens are oxidising agents.

State and explain the trend in oxidising power of the halogens.

....

....

....

.... [3]

(c) Concentrated sulfuric acid reacts with solid sodium halides.

(i) State any observations that would be made on addition of concentrated sulfuric acid to

- solid sodium chloride, ...

....

- solid sodium iodide. ....

....

[2]

(ii) Give reasons for the difference in the observations in **(i)**.

....

....

.... [2]

(iii) The addition of concentrated sulfuric acid to solid sodium bromide, NaBr, produces brown fumes and an acidic gas that decolourises acidified potassium manganate(VII) solution. This acidic gas is a significant contributor to acid rain.

Write the equation for the reaction of concentrated sulfuric acid with sodium bromide.

.... [2]

(d) An aqueous solution, **Z**, contains a mixture of sodium chloride and sodium iodide.

(i) Excess aqueous silver nitrate is added to **Z** in a test-tube. A yellow precipitate forms.

Explain the colour of this precipitate.

....

.... [1]

(ii) Aqueous ammonia is then added to the test-tube in **(i)**. The mass of precipitate decreases.

Explain this observation.

....

.... [1]

[S'17 3 Q2]

17 The elements in Group 17, the halogens, show trends in both their chemical and physical properties. The elements and their compounds have a wide variety of uses.

(a) At room temperature fluorine and chlorine are gases, bromine is a liquid and iodine is a solid.

(i) State the trend in the volatility of the Group 17 elements down the group.

.... [1]

(ii) Explain this trend.

....

....

.... [2]

(b) Iodine, I_2, can be displaced from NaI(aq), by chlorine, Cl_2.

Write an equation for this reaction.

.... [1]

(c) Silver nitrate solution, $AgNO_3$(aq), is added to separate solutions of NaI and NaC*l*. Precipitates form. An excess of aqueous ammonia is then added to both precipitates.

(i) Complete the table to give the colour and name of the precipitate formed in each reaction and the effect of the addition of an excess of aqueous ammonia to each of the precipitates formed.

	NaI(aq) + $AgNO_3$(aq)	NaC*l*(aq) + $AgNO_3$(aq)
colour of precipitate		
name of precipitate		
effect of addition of an excess of aqueous ammonia to the precipitate		

[3]

(ii) Write an ionic equation, including state symbols, to show the reaction occurring when $AgNO_3$(aq) is added to NaI(aq).

.... [1]

(d) Solid NaI reacts with concentrated sulfuric acid to form purple fumes of I_2(g) and hydrogen sulfide gas, H_2S(g).
However, when solid NaC*l* reacts with concentrated sulfuric acid the only gas produced is HC*l*(g).

Explain the difference in the reactions of concentrated sulfuric acid with NaI and with NaC*l*. Your answer should refer to the role of the sulfuric acid in each reaction.

....

....

....

....

....

.... [3]

(e) Chlorine is commonly used in water purification. When chlorine is added to water it reacts to produce a mixture of acids, one of which is chloric(I) acid, HC*l*O, a powerful oxidising agent.

(i) Explain the meaning of the term *oxidising agent*, in terms of electron transfer.

....

....

.... [1]

(ii) Suggest an equation for this reaction of chlorine with water.

.... [1]

(iii) Write an equation for the reaction of chlorine with **hot** aqueous sodium hydroxide.

Use oxidation numbers to explain why this is a redox reaction.

equation ....

....

.... [2]

[S'18 3 Q3]

DATA BOOKLET

1 Important values, constants and standards

molar gas constant	$R = 8.31\ \mathrm{J\,K^{-1}\,mol^{-1}}$
the Faraday constant	$F = 9.65 \times 10^{4}\ \mathrm{C\,mol^{-1}}$
the Avogadro constant	$L = 6.02 \times 10^{23}\ \mathrm{mol^{-1}}$
the Planck constant	$h = 6.63 \times 10^{-34}\ \mathrm{J\,s}$
speed of light in a vacuum	$c = 3.00 \times 10^{8}\ \mathrm{m\,s^{-1}}$
rest mass of proton, $^{1}_{1}\mathrm{H}$	$m_p = 1.67 \times 10^{-27}\ \mathrm{kg}$
rest mass of neutron, $^{1}_{0}\mathrm{n}$	$m_n = 1.67 \times 10^{-27}\ \mathrm{kg}$
rest mass of electron, $^{0}_{-1}\mathrm{e}$	$m_e = 9.11 \times 10^{-31}\ \mathrm{kg}$
electronic charge	$e = -1.60 \times 10^{-19}\ \mathrm{C}$
molar volume of gas	$V_m = 22.4\ \mathrm{dm^3\,mol^{-1}}$ at s.t.p $V_m = 24.0\ \mathrm{dm^3\,mol^{-1}}$ under room conditions (where s.t.p. is expressed as 101 kPa, approximately, and 273 K (0 °C))
ionic product of water	$K_w = 1.00 \times 10^{-14}\ \mathrm{mol^2\,dm^{-6}}$ (at 298 K (25 °C))
specific heat capacity of water	$= 4.18\ \mathrm{kJ\,kg^{-1}\,K^{-1}}$ $(= 4.18\ \mathrm{J\,g^{-1}\,K^{-1}})$

2 Ionisation energies (1st, 2nd, 3rd and 4th) of selected elements, in kJ mol^{-1}

	Proton number	First	Second	Third	Fourth
H	1	1310	–	–	–
He	2	2370	5250	–	–
Li	3	519	7300	11800	–
Be	4	900	1760	14800	21000
B	5	799	2420	3660	25000
C	6	1090	2350	4610	6220
N	7	1400	2860	4590	7480
O	8	1310	3390	5320	7450
F	9	1680	3370	6040	8410
Ne	10	2080	3950	6150	9290
Na	11	494	4560	6940	9540
Mg	12	736	1450	7740	10500
Al	13	577	1820	2740	11600
Si	14	786	1580	3230	4360
P	15	1060	1900	2920	4960
S	16	1000	2260	3390	4540
Cl	17	1260	2300	3850	5150
Ar	18	1520	2660	3950	5770
K	19	418	3070	4600	5860
Ca	20	590	1150	4940	6480
Sc	21	632	1240	2390	7110
Ti	22	661	1310	2720	4170
V	23	648	1370	2870	4600
Cr	24	653	1590	2990	4770
Mn	25	716	1510	3250	5190
Fe	26	762	1560	2960	5400
Co	27	757	1640	3230	5100
Ni	28	736	1750	3390	5400
Cu	29	745	1960	3350	5690
Zn	30	908	1730	3828	5980
Ga	31	577	1980	2960	6190

	Proton number	First	Second	Third	Fourth
Br	35	1140	2080	3460	4850
Rb	37	403	2632	3900	5080
Sr	38	548	1060	4120	5440
Ag	47	731	2074	3361	–
I	53	1010	1840	2040	4030
Cs	55	376	2420	3300	–
Ba	56	502	966	3390	–

3 Bond energies

3(a) Bond energies in diatomic molecules (these are exact values)

Homonuclear

Bond	Energy/kJ mol^{-1}
H–H	436
D–D	442
N≡N	944
O=O	496
P≡P	485
S=S	425
F–F	158
Cl–Cl	242
Br–Br	193
I–I	151

Heteronuclear

Bond	Energy/kJ mol^{-1}
H–F	562
H–Cl	431
H–Br	366
H–I	299
C≡O	1077

3(b) Bond energies in polyatomic molecules (these are average values)

Homonuclear

Bond	Energy/kJ mol^{-1}
C–C	350
C=C	610
C≡C	840
C$\mathrel{\overset{\cdots}{-}}$C (benzene)	520
N–N	160
N=N	410
O–O	150
Si–Si	222
P–P	200
S–S	264

Heteronuclear

Bond	Energy/kJ mol^{-1}
C–H	410
C–Cl	340
C–Br	280
C–I	240
C–N	305
C=N	610
C≡N	890
C–O	360
C=O	740
C=O in CO_2	805
N–H	390
N–Cl	310
O–H	460
Si–Cl	359
Si–H	320
Si–O (in $SiO_2(s)$)	460
Si=O (in $SiO_2(g)$)	640
P–H	320
P–Cl	330
P–O	340
P=O	540
S–H	347
S–Cl	250
S–O	360
S=O	500

4 Standard electrode potential and redox potentials, $E^\ominus$ at 298 K (25 °C)

For ease of reference, two tables are given:

(a) an extended list in alphabetical order

(b) a shorter list in decreasing order of magnitude, i.e. a redox series.

(a) $E^\ominus$ in alphabetical order

Electrode reaction	$E^\ominus$/V
$Ag^+ + e^- \rightleftharpoons Ag$	+0.80
$Al^{3+} + 3e^- \rightleftharpoons Al$	–1.66
$Ba^{2+} + 2e^- \rightleftharpoons Ba$	–2.90
$Br_2 + 2e^- \rightleftharpoons 2Br^-$	+1.07
$Ca^{2+} + 2e^- \rightleftharpoons Ca$	–2.87
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$2HOCl + 2H^+ + 2e^- \rightleftharpoons Cl_2 + 2H_2O$	+1.64
$ClO^- + H_2O + 2e^- \rightleftharpoons Cl^- + 2OH^-$	+0.89
$Co^{2+} + 2e^- \rightleftharpoons Co$	–0.28
$Co^{3+} + e^- \rightleftharpoons Co^{2+}$	+1.82
$[Co(NH_3)_6]^{2+} + 2e^- \rightleftharpoons Co + 6NH_3$	–0.43
$Cr^{2+} + 2e^- \rightleftharpoons Cr$	–0.91
$Cr^{3+} + 3e^- \rightleftharpoons Cr$	–0.74
$Cr^{3+} + e^- \rightleftharpoons Cr^{2+}$	–0.41
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33
$Cu^+ + e^- \rightleftharpoons Cu$	+0.52
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
$Cu^{2+} + e^- \rightleftharpoons Cu^+$	+0.15
$[Cu(NH_3)_4]^{2+} + 2e^- \rightleftharpoons Cu + 4NH_3$	–0.05
$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	–0.44
$Fe^{3+} + 3e^- \rightleftharpoons Fe$	–0.04
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77
$[Fe(CN)_6]^{3-} + e^- \rightleftharpoons [Fe(CN)_6]^{4-}$	+0.36
$Fe(OH)_3 + e^- \rightleftharpoons Fe(OH)_2 + OH^-$	–0.56
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$2H_2O + 2e^- \rightleftharpoons H_2 + 2OH^-$	–0.83
$I_2 + 2e^- \rightleftharpoons 2I^-$	+0.54

Electrode reaction			$E^\ominus$ / V
$K^+ + e^-$	$\rightleftharpoons$	K	–2.92
$Li^+ + e^-$	$\rightleftharpoons$	Li	–3.04
$Mg^{2+} + 2e^-$	$\rightleftharpoons$	Mg	–2.38
$Mn^{2+} + 2e^-$	$\rightleftharpoons$	Mn	–1.18
$Mn^{3+} + e^-$	$\rightleftharpoons$	Mn^{2+}	+1.49
$MnO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Mn^{2+} + 2H_2O$	+1.23
$MnO_4^- + e^-$	$\rightleftharpoons$	MnO_4^{2-}	+0.56
$MnO_4^- + 4H^+ + 3e^-$	$\rightleftharpoons$	$MnO_2 + 2H_2O$	+1.67
$MnO_4^- + 8H^+ + 5e^-$	$\rightleftharpoons$	$Mn^{2+} + 4H_2O$	+1.52
$NO_3^- + 2H^+ + e^-$	$\rightleftharpoons$	$NO_2 + H_2O$	+0.81
$NO_3^- + 3H^+ + 2e^-$	$\rightleftharpoons$	$HNO_2 + H_2O$	+0.94
$NO_3^- + 10H^+ + 8e^-$	$\rightleftharpoons$	$NH_4^+ + 3H_2O$	+0.87
$Na^+ + e^-$	$\rightleftharpoons$	Na	–2.71
$Ni^{2+} + 2e^-$	$\rightleftharpoons$	Ni	–0.25
$[Ni(NH_3)_6]^{2+} + 2e^-$	$\rightleftharpoons$	$Ni + 6NH_3$	–0.51
$H_2O_2 + 2H^+ + 2e^-$	$\rightleftharpoons$	$2H_2O$	+1.77
$HO_2^- + H_2O + 2e^-$	$\rightleftharpoons$	$3OH^-$	+0.88
$O_2 + 4H^+ + 4e^-$	$\rightleftharpoons$	$2H_2O$	+1.23
$O_2 + 2H_2O + 4e^-$	$\rightleftharpoons$	$4OH^-$	+0.40
$O_2 + 2H^+ + 2e^-$	$\rightleftharpoons$	H_2O_2	+0.68
$O_2 + H_2O + 2e^-$	$\rightleftharpoons$	$HO_2^- + OH^-$	–0.08
$Pb^{2+} + 2e^-$	$\rightleftharpoons$	Pb	–0.13
$Pb^{4+} + 2e^-$	$\rightleftharpoons$	Pb^{2+}	+1.69
$PbO_2 + 4H^+ + 2e^-$	$\rightleftharpoons$	$Pb^{2+} + 2H_2O$	+1.47
$SO_4^{2-} + 4H^+ + 2e^-$	$\rightleftharpoons$	$SO_2 + 2H_2O$	+0.17
$S_2O_8^{2-} + 2e^-$	$\rightleftharpoons$	$2SO_4^{2-}$	+2.01
$S_4O_6^{2-} + 2e^-$	$\rightleftharpoons$	$2S_2O_3^{2-}$	+0.09
$Sn^{2+} + 2e^-$	$\rightleftharpoons$	Sn	–0.14
$Sn^{4+} + 2e^-$	$\rightleftharpoons$	Sn^{2+}	+0.15
$V^{2+} + 2e^-$	$\rightleftharpoons$	V	–1.20
$V^{3+} + e^-$	$\rightleftharpoons$	V^{2+}	–0.26
$VO^{2+} + 2H^+ + e^-$	$\rightleftharpoons$	$V^{3+} + H_2O$	+0.34
$VO_2^+ + 2H^+ + e^-$	$\rightleftharpoons$	$VO^{2+} + H_2O$	+1.00
$VO_3^- + 4H^+ + e^-$	$\rightleftharpoons$	$VO^{2+} + 2H_2O$	+1.00
$Zn^{2+} + 2e^-$	$\rightleftharpoons$	Zn	–0.76

(b) $E^\ominus$ in decreasing order of oxidising power

(a selection only – see also the extended alphabetical list on the previous pages)

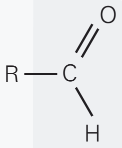
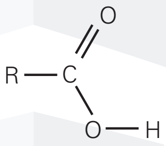
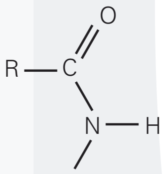
Electrode reaction	$E^\ominus$/V
$F_2 + 2e^- \rightleftharpoons 2F^-$	+2.87
$S_2O_8^{2-} + 2e^- \rightleftharpoons 2SO_4^{2-}$	+2.01
$H_2O_2 + 2H^+ + 2e^- \rightleftharpoons 2H_2O$	+1.77
$MnO_4^- + 8H^+ + 5e^- \rightleftharpoons Mn^{2+} + 4H_2O$	+1.52
$PbO_2 + 4H^+ + 2e^- \rightleftharpoons Pb^{2+} + 2H_2O$	+1.47
$Cl_2 + 2e^- \rightleftharpoons 2Cl^-$	+1.36
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightleftharpoons 2Cr^{3+} + 7H_2O$	+1.33
$O_2 + 4H^+ + 4e^- \rightleftharpoons 2H_2O$	+1.23
$Br_2 + 2e^- \rightleftharpoons 2Br^-$	+1.07
$ClO^- + H_2O + 2e^- \rightleftharpoons Cl^- + 2OH^-$	+0.89
$NO_3^- + 10H^+ + 8e^- \rightleftharpoons NH_4^+ + 3H_2O$	+0.87
$NO_3^- + 2H^+ + e^- \rightleftharpoons NO_2 + H_2O$	+0.81
$Ag^+ + e^- \rightleftharpoons Ag$	+0.80
$Fe^{3+} + e^- \rightleftharpoons Fe^{2+}$	+0.77
$I_2 + 2e^- \rightleftharpoons 2I^-$	+0.54
$O_2 + 2H_2O + 4e^- \rightleftharpoons 4OH^-$	+0.40
$Cu^{2+} + 2e^- \rightleftharpoons Cu$	+0.34
$SO_4^{2-} + 4H^+ + 2e^- \rightleftharpoons SO_2 + 2H_2O$	+0.17
$Sn^{4+} + 2e^- \rightleftharpoons Sn^{2+}$	+0.15
$S_4O_6^{2-} + 2e^- \rightleftharpoons 2S_2O_3^{2-}$	+0.09
$2H^+ + 2e^- \rightleftharpoons H_2$	0.00
$Pb^{2+} + 2e^- \rightleftharpoons Pb$	–0.13
$Sn^{2+} + 2e^- \rightleftharpoons Sn$	–0.14
$Fe^{2+} + 2e^- \rightleftharpoons Fe$	–0.44
$Zn^{2+} + 2e^- \rightleftharpoons Zn$	–0.76
$2H_2O + 2e^- \rightleftharpoons H_2 + 2OH^-$	–0.83
$V^{2+} + 2e^- \rightleftharpoons V$	–1.20
$Mg^{2+} + 2e^- \rightleftharpoons Mg$	–2.38
$Ca^{2+} + 2e^- \rightleftharpoons Ca$	–2.87
$K^+ + e^- \rightleftharpoons K$	–2.92

5 Atomic and ionic radii

(a) Period 1	**atomic/nm**		**ionic/nm**			
single covalent	H	0.037	H^+	0.208		
van der Waals	He	0.140				
(b) Period 2						
metallic	Li	0.152	Li^+	0.060		
	Be	0.112	Be^{2+}	0.031		
single covalent	B	0.080	B^{3+}	0.020		
	C	0.077	C^{4+}	0.015	C^{4-}	0.260
	N	0.074			N^{3-}	0.171
	O	0.073			O^{2-}	0.140
	F	0.072			F^-	0.136
van der Waals	Ne	0.160				
(c) Period 3						
metallic	Na	0.186	Na^+	0.095		
	Mg	0.160	Mg^{2+}	0.065		
	Al	0.143	Al^{3+}	0.050		
single covalent	Si	0.117	Si^{4+}	0.041		
	P	0.110			P^{3-}	0.212
	S	0.104			S^{2-}	0.184
	Cl	0.099			Cl^-	0.181
van der Waals	Ar	0.190				
(d) Group 2						
metallic	Be	0.112	Be^{2+}	0.031		
	Mg	0.160	Mg^{2+}	0.065		
	Ca	0.197	Ca^{2+}	0.099		
	Sr	0.215	Sr^{2+}	0.113		
	Ba	0.217	Ba^{2+}	0.135		
	Ra	0.220	Ra^{2+}	0.140		

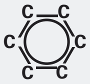
(e) Group 14	**atomic/nm**		**ionic/nm**			
single covalent	C	0.077				
	Si	0.117	Si^{4+}	0.041		
	Ge	0.122	Ge^{2+}	0.093		
metallic	Sn	0.162	Sn^{2+}	0.112		
	Pb	0.175	Pb^{2+}	0.120		
(f) Group 17						
single covalent	F	0.072	F^{-}	0.136		
	Cl	0.099	Cl^{-}	0.181		
	Br	0.114	Br^{-}	0.195		
	I	0.133	I^{-}	0.216		
	At	0.140				
(g) First row transition elements						
metallic	Sc	0.164			Sc^{3+}	0.081
	Ti	0.146	Ti^{2+}	0.090	Ti^{3+}	0.067
	V	0.135	V^{2+}	0.079	V^{3+}	0.064
	Cr	0.129	Cr^{2+}	0.073	Cr^{3+}	0.062
	Mn	0.132	Mn^{2+}	0.067	Mn^{3+}	0.062
	Fe	0.126	Fe^{2+}	0.061	Fe^{3+}	0.055
	Co	0.125	Co^{2+}	0.078	Co^{3+}	0.053
	Ni	0.124	Ni^{2+}	0.070	Ni^{3+}	0.056
	Cu	0.128	Cu^{2+}	0.073		
	Zn	0.135	Zn^{2+}	0.075		

6 Typical proton (^1H) chemical shift values (δ) relative to TMS = 0

Type of proton	Environment of proton	Example structures	Chemical shift range (δ)
C–H	alkane	$-CH_3$, $-CH_2-$, $>CH-$	0.9–1.7
	alkyl next to C=O	$CH_3-C=O$, $-CH_2-C=O$, $>CH-C=O$	2.2–3.0
	alkyl next to aromatic ring	CH_3-Ar, $-CH_2-Ar$, $>CH-Ar$	2.3–3.0
	alkyl next to electronegative atom	CH_3-O, $-CH_2-O$, $-CH_2-C\mathit{l}$, $>CH-Br$	3.2–4.0
	attached to alkyne	≡C–H	1.8–3.1
	attached to alkene	$=CH_2$, $=CH-$	4.5–6.0
	attached to aromatic ring	 –H	6.0–9.0
	aldehyde		9.3–10.5
O–H (see note below)	alcohol	RO–H	0.5–6.0
	phenol	 –OH	4.5–7.0
	carboxylic acid		9.0–13.0
N–H (see note below)	alkyl amine	R–NH–	1.0–5.0
	aryl amine	 $-NH_2$	3.0–6.0
	amide		5.0–12.0

Note: δ values for –O–H and –N–H protons can vary depending on solvent and concentration.

7 Typical carbon (^{13}C) chemical shift values (δ) relative to TMS = 0

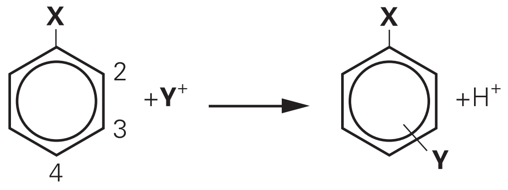
Hybridisation of the carbon atom	Environment of carbon atom	Example structures	Chemical shift range (δ)
sp^3	alkyl	$\mathbf{C}H_3$–, $\mathbf{C}H_2$–, –$\mathbf{C}$H<	0–50
sp^3	next to alkene/arene	–$\mathbf{C}H_2$–C=C, –$\mathbf{C}H_2$– 	10–40
sp^3	next to carbonyl/carboxyl	–$\mathbf{C}H_2$–COR, –$\mathbf{C}H_2$–CO_2R	25–50
sp^3	next to nitrogen	–$\mathbf{C}H_2$–NH_2, –$\mathbf{C}H_2$–NR_2, –$\mathbf{C}H_2$–NHCO	30–65
sp^3	next to chlorine (–CH_2–Br and –CH_2–I are in the same range as alkyl)	–$\mathbf{C}H_2$–C*l*	30–60
sp^3	next to oxygen	–$\mathbf{C}H_2$–OH, –$\mathbf{C}H_2$–O–CO–	50–70
sp^2	alkene or arene	>$\mathbf{C}$=$\mathbf{C}$<, 	110–160
sp^2	carboxyl	R–$\mathbf{C}O_2H$, R–$\mathbf{C}O_2R$	160–185
sp^2	carbonyl	R–$\mathbf{C}$HO, R–$\mathbf{C}$O–R	190–220
sp	alkyne	R–$\mathbf{C}$≡$\mathbf{C}$–	65–85
sp	nitrile	R–$\mathbf{C}$≡N	100–125

8 Characteristic infra-red absorption frequencies for some selected bonds

Bond	**Functional groups containing the bond**	**Absorption range (in wavenumbers)/cm^{-1}**	**Appearance of peak (*s = strong, w = weak*)**
C–O	alcohols, ethers, esters	1040–1300	***s***
C=C	aromatic compounds, alkenes	1500–1680	***w*** unless conjugated
C=O	amides ketones and aldehydes esters	1640–1690 1670–1740 1710–1750	***s*** ***s*** ***s***
C≡C	alkynes	2150–2250	***w*** unless conjugated
C≡N	nitriles	2200–2250	***w***
C–H	alkanes, CH_2–H alkenes/arenes, =C–H	2850–2950 3000–3100	***s*** ***w***
N–H	amines, amides	3300–3500	***w***
O–H	carboxylic acids, RCO_2–H H-bonded alcohol, RO–H free alcohol, RO–H	2500–3000 3200–3600 3580–3650	***s*** and very broad ***s*** ***s*** and sharp

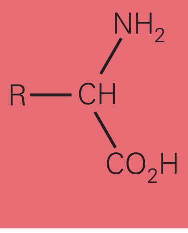
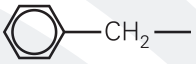
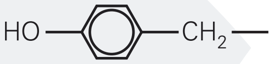
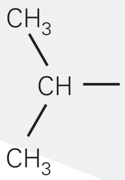
9 The orientating effect of groups in aromatic substitution reactions.

The position of the incoming group, **Y**, is determined by the nature of the group, **X**, already bonded to the ring, and not by the nature of the incoming group **Y**.



X– groups that direct the incoming Y group to the 2– or 4– positions	X– groups that direct the incoming Y group to the 3– position
$–NH_2$, –NHR or $–NR_2$	$–NO_2$
–OH or –OR	$–NH_3$
–NHCOR	–CN
$–CH_2$, –alkyl	–CHO, –COR
–C*l*	$–CO_2H$, $–CO_2R$

10 Names, structures and abbreviations of some amino acids

Name	**3-letter abbreviation**	**1-letter symbol**	**structure of side chain R– in** 
alanine	Ala	A	CH_3–
aspartic acid	Asp	D	HO_2CCH_2–
cysteine	Cys	C	$HSCH_2$–
glutamic acid	Glu	E	$HO_2CCH_2CH_2$–
glycine	Gly	G	H–
lysine	Lys	K	$H_2NCH_2CH_2CH_2CH_2$–
phenylalanine	Phe	F	
serine	Ser	S	$HOCH_2$–
tyrosine	Tyr	Y	
valine	Val	V	

Group																	
1	2											13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
3 Li lithium 6.9	4 Be beryllium 9.0			**Key** atomic number atomic symbol name relative atomic mass								5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2
11 Na sodium 23.0	12 Mg magnesium 24.3	3	4	5	6	7	8	9	10	11	12	13 A*l* aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	16 S sulfur 32.1	17 C*l* chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium –	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3
55 Cs caesium 132.9	56 Ba barium 137.3	57–71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 T*l* thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium –	85 At astatine –	86 Rn radon –
87 Fr francium –	88 Ra radium –	89–103 actinoids	104 Rf rutherfordium –	105 Db dubnium –	106 Sg seaborgium –	107 Bh bohrium –	108 Hs hassium –	109 Mt meitnerium –	110 Ds darmstadtium –	111 Rg roentgenium –	112 Cn copernicium –		114 F*l* flerovium –		116 Lv livermorium –		

lanthanoids	57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.4	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium 173.1	71 Lu lutetium 175.0
actinoids	89 Ac actinium –	90 Th thorium 232.0	91 Pa protactinium 231.0	92 U uranium 238.0	93 Np neptunium –	94 Pu plutonium –	95 Am americium –	96 Cm curium –	97 Bk berkelium –	98 Cf californium –	99 Es einsteinium –	100 Fm fermium –	101 Md mendelevium –	102 No nobelium –	103 Lr lawrencium –