Answers

Section 1 — The Structure of the Atom

Page 1 — Atomic Structure

1 Protons and neutrons.

2 +2

3 −2

Page 2 — Atomic Number, Mass Number and Isotopes

1 $A - Z = 31 - 15 =$ **16**

2 Two isotopes of the same element have the same number of protons and electrons but different numbers of neutrons.

3 All three isotopes have 6 protons and 6 electrons. Carbon-12 has (12 − 6) = 6 neutrons, carbon-13 has (13 − 6) = 7 neutrons and carbon-14 has (14 − 6) = 8 neutrons.

Page 3 — Relative Atomic Mass

1 $[(8 \times 6) + (92 \times 7)] \div 100 = \mathbf{6.92}$

2 $[(99 \times 12) + (1 \times 13)] \div 100 = \mathbf{12.01}$

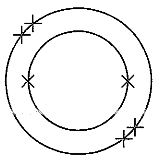
3 $[(52 \times 107) + (48 \times 109)] \div 100 = \mathbf{107.96}$

4 $23.0 + 19.0 = \mathbf{42.0}$

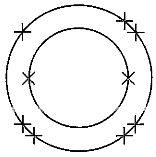
5 $12.0 + (3 \times 1.0) + 35.5 = \mathbf{50.5}$

Page 4 — Electronic Structure

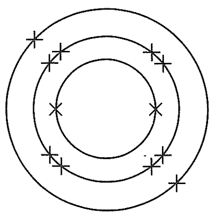
1



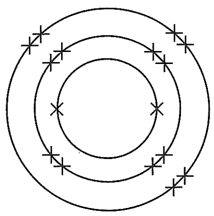
carbon



fluorine



magnesium



sulfur

2 Lithium: 2,1 Sodium: 2,8,1
Potassium: 2,8,8,1 Beryllium: 2,2
Magnesium: 2,8,2 Calcium: 2,8,8,2

3 Oxygen: $1s^2\ 2s^2\ 2p^4$
Chlorine: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$

Page 5 — The Periodic Table

1

s-block elements	p-block elements
caesium	phosphorus
potassium	aluminium
calcium	sulfur
barium	

2 E.g. They have the same number of electrons in their outer shell. / They react in similar ways.

Section 2 — Formation of Ions

Page 6 — Ionisation Energy

1 $Na_{(g)} \rightarrow Na^+_{(g)} + e^-$

2 Nuclear charge, the distance of the electron from the nucleus and shielding by inner electrons.

3 Magnesium
Fluorine
Oxygen

Page 7 — Formation of Ions

1 +1

2 Group 7

3 SO_3^{2-}

Page 8 — Oxidation Numbers

1 The oxidation number tells you how many electrons an atom has donated or accepted.

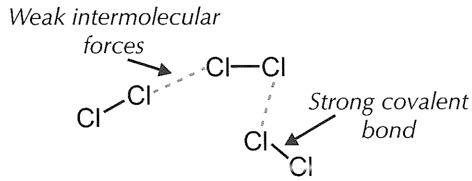
2 Al^{3+}: +3 NH_4^+: +1 Ne: 0 O^{2-}: −2

3 0

Section 3 — Intermolecular Bonding

Page 9 — Intermolecular Bonding

1

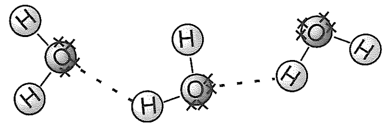


2 The trend should show an increase in boiling points as size of the alkane increases.
E.g. pentane: 36 °C, hexane: 69 °C, heptane: 98 °C, octane: 126 °C.

Page 10 — Polarity

1 a) HF as it is polar, and H_2 is non-polar.
b) H_2O as it can form hydrogen bonds, and H_2S can't.
c) CH_3F as fluorine is much more electronegative than carbon so it will be a polar molecule. Iodine is less electronegative / iodine and carbon have similar electronegativities, so CH_3I is non-polar.

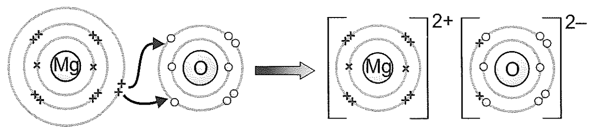
2 E.g.



Section 4 — Bonding and Properties

Page 11 — Ionic Bonding

1



2 You need two K^+ ions (2 × +1) to balance out each O^{2-} ion (1 × −2), so the ratio is 2:1. The ionic formula is K_2O.

Answers

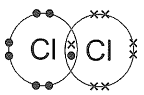
Page 12 — Ionic Compounds

1 BeO, Li_2O, LiF
The higher the charges on the ions, the stronger the bonds between them. The stronger the bonds, the higher the melting point. Beryllium oxide is formed from ions which both have charges of magnitude 2. Lithium oxide is formed from oxide ions with a –2 charge and lithium ions with only a +1 charge. The ions in lithium fluoride both have charges of magnitude one. Therefore the strongest bonds will be in beryllium oxide, followed by lithium oxide, then lithium fluoride.

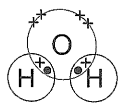
2 When potassium chloride is solid, the K^+ and Cl^- ions are held together in an ionic lattice, so they're not free to move and conduct electricity. When molten or dissolved, the ions separate, so they're free to move and able to carry a current.

Page 13 — Covalent Bonding

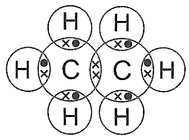
1 a) Chlorine:



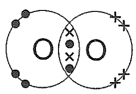
b) Water:



c) Ethane:

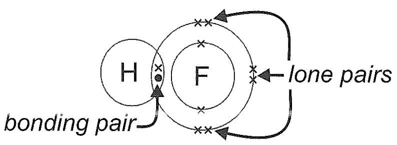


d) Oxygen:



Page 14 — Small Covalent Molecules

1



2 For a small covalent compound to boil, you have to break the weak intermolecular bonds between the molecules rather than the strong covalent bonds between the atoms in a molecule. These intermolecular bonds don't need much energy to break, so nitrogen has a low boiling point and is a gas at room temperature.

Page 15 — Giant Covalent Molecules

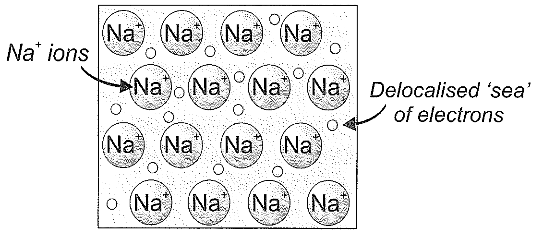
1 You could use solubility, though not all ionic compounds are soluble in water so this test may prove inconclusive. Melt and check to see if the molten substance conducts electricity— if it does it's probably ionic. You need to be careful of graphite though, which is a giant covalent molecule that is able to conduct electricity. To get around this, you could test the conductivity of the crystal in its solid form as well. While solid graphite will conduct, ionic salts only conduct electricity when molten or in solution.

2 In sodium chloride, the energy required to break the strong ionic bonds is provided when the ions become surrounded by water molecules. Fewer strong bonds are replaced by many more weaker bonds. In the case of a giant covalent molecule there is no way to get the energy required to break the many strong covalent bonds between atoms, so diamond doesn't dissolve.

Page 16 — Metallic Bonding

1 Calcium is likely to have a higher melting point. This is because calcium is made up of a lattice of Ca^{2+} ions, with two delocalised electrons per ion. Potassium however is made up of K^+ ions and only one delocalised electron per ion. So the bonding in calcium is stronger, and will require more energy to be broken leading to a higher melting point.

2



3 Similarities, e.g. high melting and boiling points / both conduct electricity when molten.
Differences, e.g. solid sodium chloride is soluble in water, sodium is not / solid sodium chloride is an electrical insulator but solid sodium conducts electricity.

Page 17 — Trends in Properties Across the Periodic Table

1 The melting points of the oxides of sodium, magnesium and aluminium are all high. Also, all three are conductors of electricity when molten. These properties clearly point towards the oxides being ionic.
Silicon dioxide also has a high melting point but is a non-conductor when molten so it has a giant covalent structure.
Oxides of phosphorus and sulfur have low melting points and are non-conductors. These are therefore likely to be small covalent molecules.

2 The melting points will be high up to silicon (which forms strong covalent bonds) then drop down to phosphorus and sulfur which are small covalent molecules.

3 a) Sodium chloride has ionic bonds.
b) The chloride of phosphorus has covalent bonds and is a small covalent molecules.

Section 5 — Chemical Equations

Page 19 — Writing and Balancing Equations

1 Step 2: $CH_4 + O_2 \rightarrow CO_2 + H_2O$
Step 3: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$
(The Cs already balance, so put a 2 in front of H_2O to balance the Hs. Now put a 2 in front of O_2 to balance the Os. Check that all still balances.)

2 a) $C_2H_5OH + 3O_2 \rightarrow 2CO_2 + 3H_2O$
b) $Ca(OH)_2 + 2HCl \rightarrow CaCl_2 + 2H_2O$

3 First balance the atoms:
$Cl_{2(g)} + Fe^{2+}_{(aq)} \rightarrow 2Cl^-_{(aq)} + Fe^{3+}_{(aq)}$
Then balance the charges:
$Cl_{2(g)} + 2Fe^{2+}_{(aq)} \rightarrow 2Cl^-_{(aq)} + 2Fe^{3+}_{(aq)}$

Answers

Section 6 — Inorganic Chemistry

Page 20 — Group 2

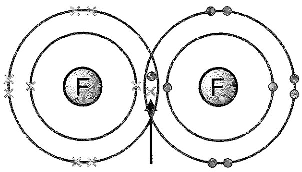
1 *Reactivity increases down Group 2, so you'd expect magnesium to react very slowly with cold water to produce hydrogen gas and magnesium hydroxide. You'd expect strontium to react vigorously with cold water to produce hydrogen gas and strontium hydroxide.*

2 *The boiling points of the Group 2 metals will decrease down the group. This is because, as you go down the Group, the nuclei become more shielded. This causes the attraction between the positive metal ions and the free electrons in the metal to decrease. So the strength of the metallic bonds decreases down the group, making them easier to break.*

Page 21 — Group 7

1 a) *Chlorine is more reactive than bromine, so it displaces the bromide in solution producing chloride and bromine.*
b) *Iodine is less reactive than chlorine, so no reaction takes place.*
c) *Iodine is less reactive than bromine, so no reaction takes place.*
d) *Chlorine is more reactive than iodine, so it displaces the iodide in solution producing chloride and iodine.*

2



Page 22 — Acids and Bases

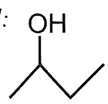
1 $HNO_3 + KOH \rightarrow KNO_3 + H_2O$

2 a) $H_2SO_4 \xrightarrow{water} 2H^+ + SO_4^{2-}$
b) $KOH \xrightarrow{water} K^+ + OH^-$
c) $HNO_3 \xrightarrow{water} H^+ + NO_3^-$

Section 7 — Organic Chemistry

Page 23 — Organic Molecules

1 skeletal:



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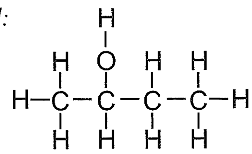
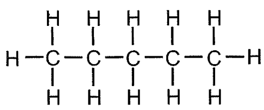


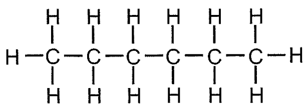
2 $C_3H_6O_2$

Page 24 — Alkanes

1 pentane:



hexane:

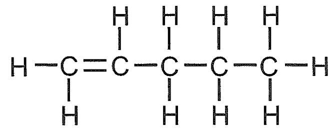


2 a) CH_4, C_2H_6, C_3H_8, C_4H_{10}.
b) General formula: C_nH_{2n+2}

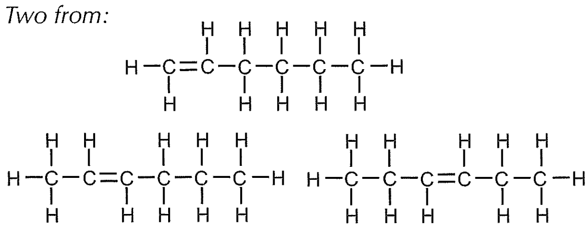
3 $CH_3CH_2CH_{3\,(g)} + 5O_{2\,(g)} \rightarrow 3CO_{2\,(g)} + 4H_2O_{(g)}$

Page 25 — Alkenes

1 e.g.



2 Two from:

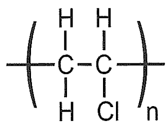


3 General formula: C_nH_{2n}

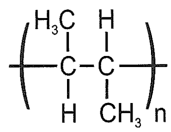
Page 26 — Polymerisation

1 *Alkenes have a double bond that can open and link to other monomers.*

2 a)

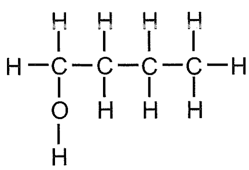


b)

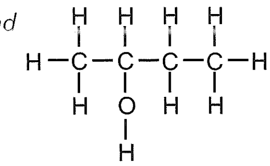


Page 27 — Alcohols

1



and



2 General formula: $C_nH_{(2n+1)}OH$

3 *Ethanol will have a higher melting point than ethane because it is able to form hydrogen bonds. These require more energy to break than the intermolecular bonds that form between nonpolar molecules such as ethane.*

Section 8 — Chemical Reactions

Page 30 — Reaction Types

1 a) combustion, exothermic, oxidation, redox, (reduction)
b) displacement, oxidation, precipitation, redox, reduction, substitution
c) exothermic, neutralisation, oxidation, substitution, redox, reduction
d) addition, hydrogenation, (oxidation, redox, reduction)

Answers

Section 9 — Rates of Reaction

Page 31 — Reaction Rates

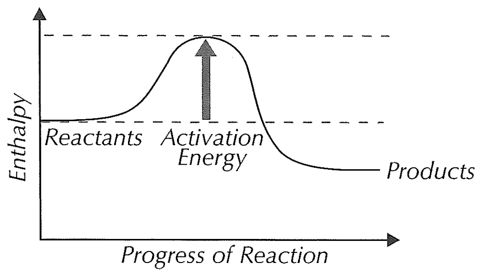
1 a) e.g. Measure the change in temperature / Measure the change in pH / Measure the loss of mass as CO_2 is evolved / Measure the volume of CO_2 produced using a gas syringe.
b) e.g. Measure how long it takes for the solution to go cloudy / Measure the loss of mass as SO_2 is evolved / Measure the volume of SO_2 produced using a gas syringe.
c) e.g. Measure the change in pH / Measure the loss of mass as CO_2 is evolved / Measure the volume of CO_2 produced using a gas syringe.

Page 32 — Collision Theory

1 The particles may have collided facing in the wrong direction or without enough energy to react.

2 The activation energy is the minimum amount of kinetic energy particles need to react.

3



Page 33 — Reaction Rates and Catalysts

1 e.g. Increase the temperature / Increase the concentration of the reactants / Add a catalyst.

2 The particles in the gas will be closer together so they're more likely to collide.

3 A catalyst is a substance that speeds up the rate of reaction without being changed or used up itself.

4 E.g. Catalysts make reactions cheaper to run and reduce their CO_2 emissions.

Section 10 — Equilibria

Page 34 — Reversible Reactions

1 a) At the start of the reaction, the rate of the forward reaction is faster than the rate of the backwards reaction.
b) At equilibrium, the rates of the forward and backward reactions are the same.

2 Dynamic equilibrium is where the forwards and backwards reactions of a reversible reaction are going at the same rate. This means that although both reactions are still happening, the concentrations of the products and reactants don't change.

Page 35 — Le Chatelier's Principle

1 Increasing the amount of steam will increase the concentration of particles on the left of the equation (which will also increase the pressure on the left hand side), and move the position of equilibrium to the right, increasing the yield of ethanol.

2 a) The equilibrium will move to the left to favour the endothermic reaction.
b) The equilibrium will move to the right to try and increase the concentration of ammonia.

Page 36 — Equilibrium and Yield

1 The temperature is low, which would favour the forward reaction, and increase the yield of ethanol.
But it is so low that the forward reaction rate will be much too slow to be economic.

2 The pressure is high, which would favour the forward reaction, and increase the yield of ethanol.
But such a high pressure would be very expensive to maintain, making the reaction uneconomic.

Section 11 — Calculations

Page 37 — The Mole

1 molar mass = mass ÷ moles
= 68.6 ÷ 0.700 = **98.0 g mol^{-1}**

2 Molar mass NaCl = 23.0 + 35.5 = 58.5 g mol^{-1}
moles = mass ÷ molar mass = 117 ÷ 58.5 = **2.00 moles**

3 Molar mass water = 16.0 + (2 × 1.0) = 18.0 g mol^{-1}
moles of water = 54.0 ÷ 18.0 = 3.00 moles.
Molar mass of iron = 55.8 g mol^{-1}
moles of iron = 84.0 ÷ 55.8 = 1.51 moles.
There are more moles of water.

Page 38 — Determination of Formulae from Experiments

1 a) Mass of each substance: Pb: 12.9 g O: 1.00 g
Number of moles of each substance:
Pb: 12.9 ÷ 207.2 = 0.0623 moles
O: 1.00 ÷ 16.0 = 0.0625 moles
Divide by the smallest number (0.0623):
Pb: 1.00 O: 1.00
The ratio of Pb : O is 1 : 1. The empirical formula is **PbO**.
b) Mass of each substance: Fe: 2.33 g O: 1.00 g
Number of moles of each substance:
Fe: 2.33 ÷ 55.8 = 0.0418 moles
O: 1.00 ÷ 16 = 0.0625 moles
Divide by the smallest number (0.0418):
Fe: 1.00 O: 1.50
Multiply by two to give whole numbers: Fe: 2 O: 3
The ratio of Fe : O is 2 : 3. The empirical formula is $\mathbf{Fe_2O_3}$.

2 Percentage composition of each substance:
H: 4.30% C: 26.1% O: 69.6%
Number of moles of each substance:
H: 4.3 ÷ 1.0 = 4.30 moles
C: 26.1 ÷ 12.0 = 2.18 moles
O: 69.6 ÷ 16.0 = 4.35 moles
Divide by the smallest number (2.18):
H: 1.97 C: 1.00 O: 2.00
The ratio of H : C : O is 2 : 1 : 2.
The empirical formula is $\mathbf{CH_2O_2}$.

Answers

Page 39 — Calculation of Molecular Formulae

1 Mass of each substance in 100 g:
C: 52.2 g H: 13.0 g O: 34.8 g
Number of moles of each substance:
C: 52.2 ÷ 12.0 = 4.35 moles
H: 13.0 ÷ 1.0 = 13.0 moles
O: 34.8 ÷ 16.0 = 2.18 moles
Divide by the smallest number (2.18):
C: 2.00 H: 5.96 O: 1.00
Ratio of C : H : O is 2 : 6 : 1.
The empirical formula is C_2H_6O.
The empirical formula has a relative mass of
(2 × 12.0) + (6 × 1.0) + (1 × 16.0) = 46.0. This is the same as the relative formula mass of the compound, so the molecular formula is also $\mathbf{C_2H_6O}$.

2 Mass of each substance in 100 g:
C: 92.3 g H: 7.70 g
Number of moles of each substance:
C: 92.3 ÷ 12.0 = 7.69 moles
H: 7.70 ÷ 1.0 = 7.70 moles
Divide by the smallest number (7.69):
C: 1.00 H: 1.00
So the ratio of C : H is 1 : 1. The empirical formula is CH.
The empirical formula has a relative formula mass of
12.0 + 1.0 = 13.0.
78.0 ÷ 13.0 = 6.00, so there are six lots of the empirical formula in the compound.
The molecular formula is $\mathbf{C_6H_6}$.

3 a) $\frac{1 \times 16.0}{(2 \times 12.0) + (6 \times 1.0) + (1 \times 16.0)} \times 100 = \mathbf{34.8\%}$

b) $\frac{3 \times 16.0}{(1 \times 1.0) + (1 \times 14.0) + (3 \times 16.0)} \times 100 = \mathbf{76.2\%}$

c) $\frac{1 \times 16.0}{(3 \times 12.0) + (6 \times 1.0) + (1 \times 16.0)} \times 100 = \mathbf{27.5\%}$

Page 40 — Atom Economy

1 a) M_r C_2H_5OH = (2 × 12.0) + (6 × 1.0) + 16.0 = 46.0
M_r NaBr = 23.0 + 79.9 = 102.9
$\frac{46.0}{46.0 + 102.9} \times 100 = \mathbf{30.9\%}$

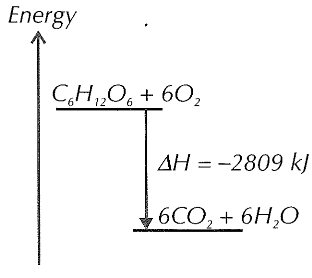
b) This reaction has an atom economy of 100%, so will be cheaper and greener to run.

Section 12 — Enthalpy

Page 41 — Endothermic and Exothermic Reactions

1 a) exothermic
b) endothermic
c) exothermic
d) endothermic

2 a)



b) exothermic

Page 42 — Bond Energy

1 a) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms:
8 C — H bonds = 8 × 413 = 3304
2 C — C bonds = 2 × 348 = 696
5 O = O bonds = 5 × 498 = 2490
TOTAL = 6490
Step 2: Calculate the energy released by all the new bonds formed between the product atoms:
6 C = O bonds = 6 × 743 = 4458
8 O — H bonds = 8 × 463 = 3704
TOTAL = 8162
Step 3: Find the overall value for the energy change:
+6490 – 8162 = **–1672 kJ mol^{-1}**

b) Step 1: Calculate the energy required to break all of the bonds between the reactant atoms:
1 C — C bond = 1 × 348 = 348
1 C — O bond = 1 × 360 = 360
5 C — H bonds = 5 × 413 = 2065
1 O — H bonds = 1 × 463 = 463
3 O = O bond = 3 × 498 = 1494
TOTAL = 4730
Step 2: Calculate the energy released by all the new bonds formed between the product atoms
4 C = O bonds = 4 × 743 = 2972
6 O — H bonds = 6 × 463 = 2778
TOTAL = 5750
Step 3: Find the overall value for the energy change:
+4730 – 5750 = **–1020 kJ mol^{-1}**

c) Step 1: Calculate the energy required to break the H–H and C=C bonds:
1 H — H bond = 1 × 436 = 436
1 C = C bond = 1 × 612 = 612
TOTAL = 1048
Step 2: Calculate the energy released by all the new bonds formed between product atoms
2 C — H bonds = 2 × 413 = 826
1 C — C bond = 1 × 348 = 348
TOTAL = 1174
Step 3: Find the overall value for the energy change:
+1048 – 1174 = **–126 kJ mol^{-1}**

Section 13 — Investigating & Interpreting

Page 43 — Planning Experiments

1 a) The time taken for the lump of magnesium to react.
b) E.g. The mass of magnesium used / The concentration of hydrochloric acid / The volume of hydrochloric acid / The surface area of the magnesium.

Page 44 — Presenting and Interpreting Data

1 The anomalous result is 19.0 cm^3.
The mean is: (22.0 + 23.0 + 22.0 + 24.0) ÷ 4 = **22.8 cm^3**

Page 45 — Conclusions and Error

1 a) The uncertainty of the weighing scales is 0.05 g, so the percentage error = $\frac{0.05}{1.4} \times 100 = \mathbf{3.6\%}$

b) The uncertainty of the clock is 0.5 s, so the percentage error = $\frac{0.5}{23} \times 100 = \mathbf{2.2\%}$

c) The uncertainty of the thermometer is 0.1 °C, so the percentage error = $\frac{0.1}{10.6} \times 100 = \mathbf{0.9\%}$

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