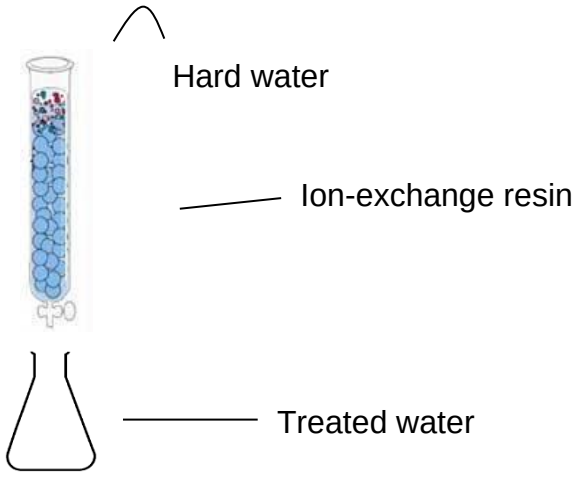
1) Atoms, Moles and Stoichiometry

1 The hardness present in a water sample due to dissolved calcium ions can be determined by using ion-exchange column as shown in the diagram.

$$Ca^{2+} + 2H^{+}\text{-resin} \rightarrow Ca^{2+}\text{-resin} + 2H^{+}$$

A 50 cm^3 sample of a solution containing calcium sulfate was passed through the ion-exchange resin. The calcium ions in the sample were quantitatively exchanged by hydrogen ions. The sample collected in the flask required 25 cm^3 of 1.0×10^{-2} mol dm^{-3} potassium hydroxide for complete neutralisation.



What was the concentration of calcium sulfate in the original sample?

A 2.5×10^{-3} mol dm^{-3} **B** 5.0×10^{-3} mol dm^{-3}

C 2.0×10^{-2} mol dm^{-3} **D** 4.0×10^{-2} mol dm^{-3}

2 5.00 g of a fertilizer containing ammonium sulfate was dissolved in water and diluted to 50 cm^3. 25.0 cm^3 of this solution was added to an excess of sodium hydroxide solution. The reaction between ammonium sulfate and NaOH(aq) is given by the following equation:

$$(NH_4)_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2\ H_2O + 2\ NH_3$$

The ammonia produced was passed into 50.0 cm^3 of 0.050 mol dm^{-3} sulfuric acid solution. The residual acid then required 25.80 cm^3 of 0.100 mol dm^{3} sodium hydroxide solution for complete neutralisation.

What is the percentage by mass of ammonium sulfate (M_r = 132.1) in the fertiliser?

A 10.6% **B** 16.0% **C** 32.0% **D** 40.0%

3 A student prepared a 0.100 mol dm^{-3} solution of $Ba(OH)_2.8H_2O$ and left it in a beaker. A week later, he returned to the laboratory, used the solution for titration with 0.100 mol dm^{-3} HC*l* and was surprised to discover his titres were lower than expected.

Which of the following explains why the volume of HC*l* used was lower than expected?

A Some water had evaporated from the barium hydroxide solution

B The concentration of HC*l* was less than the stated 0.100 mol dm^{-3}.

C The barium hydroxide crystals had less water of crystallisation than stated.

D Some of the barium hydroxide had reacted with carbon dioxide in the air to form solid barium carbonate.

4 *Use of the Data Booklet is relevant to this question.*

How many oxygen atoms are present in 24 cm^3 of oxygen gas at r.t.p.?

A $24 \times 6.02 \times 10^{23} \times 32$ **B** $\frac{2 \times 24 \times 24000}{6.02 \times 10^{23}}$

C $\frac{2 \times 6.02 \times 10^{23}}{24000}$ **D** $\frac{2 \times 6.02 \times 10^{23}}{1000}$

5 Mixtures of argon and another gas are commonly used during welding. One such gaseous mixture has a density of 1.82 g dm^{-3} at s.t.p.

What could be the other gas in this mixture? [Density of Ar = 1.78 g dm^{-3} at s.t.p.]

A Carbon Dioxide **B** Oxygen **C** Nitrogen **D** Helium

6 **X** is an organic iron compound containing only Fe, C and H. When a 0.944 g sample of **X** was subjected to complete combustion, 2.23 g of CO_2 and 0.457 g of H_2O were formed.

What is the empirical formula of **X**?

A FeCH **B** FeC_2H **C** $Fe_2C_5H_2$ **D** $FeC_{10}H_{10}$

7 0.0337 g of an organic compound when vapourised completely at 100 kPa and 300 K occupied a volume of 20 cm^3. When this volume of vapour was completely burnt in excess oxygen, 40 cm^3 of carbon dioxide and 20 cm^3 of water vapour were formed. All volumes were measured at the same temperature and pressure.

What is the formula of the organic compound?

A C_2H_2 **B** C_3H_6 **C** CH_2CO **D** CH_3CHO

8 A sample of 10 dm^3 of polluted air at r.t.p is passed through limewater so that all carbon dioxide present is precipitated as calcium carbonate. The mass of calcium carbonate formed is 0.05 g.

What is the percentage, by volume, of carbon dioxide in the air sample? [A_r **C**, 12.0; O, 16.0; Ca, 40.1]

A 0.03% **B** 0.05% **C** 0.12% **D** 0.60%

9 A student added 10 cm^3 of 0.2 mol dm^{-3} KI(aq) to 10 cm^3 of 0.2 mol dm^{-3} $CuSO_4$(aq) in a beaker and observed that a white precipitate in brown solution was obtained. She knew that if she were to add $S_2O_3^{2-}$(aq) to the mixture, the brown solution will be decolourised.

Calculate the volume of 0.04 mol dm^{-3} $S_2O_3^{2-}$(aq) she should add to the mixture in the beaker in order to completely decolourise the brown solution.

A 12.5 cm^3 **B** 15.0 cm^3 **C** 16.7 cm^3 **D** 25.0 cm^3

10 Which of the following contains the greatest number of moles of particles?

A 600 g of I_2(s)

B 50 cm^3 of $H_2O(l)$

C 50 dm^3 of HC*l*(g) at s.t.p.

D 550 cm^3 of 2.5 mol dm^{-3} CH_3CO_2H(aq)

11 Naturally-occurring silicon is a mixture of three isotopes, ^{28}Si, ^{29}Si and ^{30}Si.

The relative atomic mass of silicon is 28.109

What could be the relative abundance of each of the three isotopes?

A 91.1% ^{28}Si, 7.9% ^{29}Si and 1.0% ^{30}Si

B 92.2% ^{28}Si, 4.7% ^{29}Si and 3.1% ^{30}Si

C 95.0% ^{28}Si, 3.2% ^{29}Si and 1.8% ^{30}Si

D 96.3% ^{28}Si, 0.3% ^{29}Si and 3.4% ^{30}Si

12 Which statements about 28.0 g sample of $^{14}N_2$ are correct?

1 The number of atoms is the same as the number of atoms in 24.0 g of ^{12}C.

2 The number of atoms is the same as the number of atoms in 4.0 g of ^{4}He.

3 The number of molecules is the same as the number of atoms in 32.0 g of $^{16}O_2$.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

13 Which statement about one mole of a metal is always correct?

A It contains the same number of atoms as $\frac{1}{2}$ mol of hydrogen gas

B It contains the same number of atoms as $\frac{1}{12}$ mol of ^{12}C.

C It has the same mass as 1 mol of hydrogen atoms.

D It is liberated by 1 mol of electrons.

14 Naturally occurring iron consists of four isotopes, ^{54}Fe, ^{56}Fe, ^{57}Fe and ^{58}Fe.
Two ions formed by the isotopes have chemical formulae shown below:

$$^{54}_{26}Fe^{2+} \qquad ^{56}_{26}Fe^{2+}$$

Which of the following statements is true?

A The electron configuration of both Fe^{2+} ions is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^4 4s^2$.

B When subjected to an electric field, $^{56}_{26}Fe^{2+}$ ion gives a greater angle of deflection than $^{54}_{26}Fe^{2+}$ ion

C $^{56}_{26}Fe^{2+}$ ion has more protons in its nucleus than $^{54}_{26}Fe^{2+}$ ion.

D The total number of protons and electrons in each of the Fe^{2+} ions is smaller than the nucleon number of the respective ion.

15 (SAJC 14) A solid hydrocarbon was completely combusted in a closed vessel at 120 ^{o}C. The residual gas had a volume of 64 cm^3, which decreased by 24 cm^3 after bubbling through a dehydrating agent. After this, 40% of the final gas volume consisted of oxygen. What is the empirical formula of the hydrocarbon?

A CH **B** CH_2 **C** CH_3 **D** C_2H_3

16 (NYJC 14) 0.02 mol of $C_xH_yNO_2$ was burned in excess of oxygen as shown by the equation below.

$$C_xH_yNO_2 + aO_2 \rightarrow xCO_2 + NO_2 + \frac{y}{2} H_2O$$

The gases that were produced were first passed through a U-tube containing phosphorus pentoxide and then bubbled through 0.25 mol of NaOH(aq).

The equations illustrating the reactions of the gases with NaOH(aq) are as shown below.

$$CO_2 + 2NaOH \rightarrow Na_2CO_3 + H_2O$$

$$2NO_2 + 2NaOH \rightarrow NaNO_2 + NaNO_3 + H_2O$$

The phosphorus pentoxide U-tube increased in mass by 1.15 g and the excess NaOH required 40.00 cm^3 of 0.875 mol dm^{-3} H_2SO_4 for complete reaction.

What is the ratio of x:y?

A 2:3
B 4:3
C 4:7
D 8:7

17 (VJC 13) 25.0 cm^3 of a solution of hydrogen peroxide, H_2O_2, was oxidised by excess potassium manganate(VII) solution under acidic condition. The volume of oxygen gas produced at room temperature and pressure was 26 cm^3.
What is the concentration of hydrogen peroxide in the solution?

A 2.89×10^{-2} mol dm^{-3}
B 4.33×10^{-2} mol dm^{-3}
C 8.66×10^{-2} mol dm^{-3}
D 1.30×10^{-1} mol dm^{-3}

Section B

1(a) [CJC/07]

0.50 g of hydrated iron(II) sulfate ($FeSO_4.7H_2O$) was dissolved in dilute H_2SO_4 and titrated with 0.020 mol dm^{-3} potassium manganate(VII).

What is the volume of potassium manganate(VII) required to complete the titration? [3]

1 (b) An ore contains 0.77% by mass of the mineral Cu_2S and no other copper minerals. Calculate the maximum mass of pure copper that could be produced from 5 tonnes of the ore. [1 tonne = 1000 kg] [2]

2 [AJC/08]

Industrially, 1,2-dichloroethane is manufactured through the oxychlorination of ethene. A mixture of dry HCl and pure oxygen react with ethene in a reaction chamber maintained at 350 °C to form 1,2-dichloroethane and water.

$$C_2H_4 + 2HCl + \frac{1}{2}O_2 \rightarrow C_2H_4Cl_2 + H_2O$$

To measure the percentage conversion of ethene to 1,2-dichloroethane, the amount of unreacted HCl that emerges from the chamber is bubbled through 1 dm^3 of

0.100 mol dm^{-3} NaOH for one minute. The resultant mixture is then analysed by a titrimetric method.

In one investigation, the feed gas entering the reaction chamber contains $C_2H_4(g)$, HCl(g) and $O_2(g)$ in molar ratio 1: 2: ½ . The flow rate of HCl into the reaction chamber is 3.65 g min^{-1}.

(i) Calculate the flow rate of ethene, in mol min^{-1}, into the reaction chamber?

(ii) In the titrimetric analysis, a 10 cm^3 sample of the resultant mixture requires exactly 2.50 cm^3 of a 0.100 mol dm^{-3} HNO_3 solution for complete neutralisation.
Calculate the number of moles of unreacted NaOH present in the 1 dm^3 solution.

(iii) Calculate the number of moles of unreacted HC*l* that emerged from the reaction chamber in one minute.

(iv) Hence, calculate the molar percentage conversion per minute of ethene to 1,2-dichloroethane in the oxychlorination process.

(v) State one assumption made in your calculations.

[Total: 6]

3 [TPJC/08]

Calcium ion, Ca^{2+}, is required for blood clotting and other cell processes. An abnormal Ca^{2+} concentration in the body is indicative of disease.

To measure Ca^{2+} concentration, 1 cm^3 of human blood was treated with sodium ethanedioate($Na_2C_2O_4$) solution. The resulting CaC_2O_4 precipitate was filtered and dissolved in dilute acid. This solution then required 20.50 cm^3 of 4.88 x 10^{-5} mol dm^{-3} $KMnO_4$ for complete reaction.

(a) The dark purple colour of the potassium manganate(VII) makes it suitable for its use as described above.

State the formula of the manganese-containing species that would be formed in the reaction described above. Hence, suggest the colour changes that would be observed. [2]

(b) Carbon dioxide is formed in the reaction between ethanedioate and manganate(VII) ions. Use your answer to part **(a)**, together with the *Data Booklet* to write an equation for the reaction. [2]

(c) Determine the amount of Ca^{2+} ions present in the blood sample, and hence, the concentration of Ca^{2+} in terms of mg per 100 m*l* of blood. [1 m*l* = 1 cm^3]. [3]

4 Naturally occurring gallium, Ga,(relative atomic mass: 69.7) is a mixture of two isotopes, gallium-69 and gallium-71.

(i) What do you understand by the term *relative atomic mass*? [2]

Using the information given above, calculate the percentage abundance of each isotope.

[3]

Answers: 1) Atoms, Molecules and Stoichiometry

1	A	2	C	3	D	4	D	5	A
6	D	7	C	8	C	9	D	10	B
11	B	12	D	13	A	14	D	15	B
16	A	17	B						

2) Redox

MCQ

1 A 25 cm^3 sample of a 0.1 mol dm^{-3} **B**Cl_2 solution requires 10 cm^3 of 0.10 mol dm^{-3} of acidified potassium manganate(VII) solution for complete reaction.

What is the final oxidation state of **B**?

A +3 **B** +4 **C** +5 **D** +6

2 In a historically famous experiment Wöhler 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^-

3 0.5 g of zinc powder was found to reduce an acidified solution of 25.50 cm^3 of 0.200 mol dm^{-3} metal **A** ion, **A**O_2^+. What is the final oxidation number of metal **A**?

A 0 **B** +1 **C** +2 **D** +3

4 An element **Y** can exist in a few oxidation states. 0.01 mol of Y^{2+} requires 0.004 mol of acidified $KMnO_4$ for complete reaction. The half-equation for reduction of MnO_4^- is:

$MnO_4^- + 8H^+ + 5e \rightleftharpoons Mn^{2+} + 4H_2O$

What is the final oxidation state of **Y**?

A +4 **B** +3 **C** 0 **D** −1

5 In an experiment, 100 cm^3 of 0.020 mol dm^{-3} acidified iron(II) sulfate is quantitatively by 50 cm^3 of 0.020 mol dm^{-3} of an oxidising agent, XO_3^-.

What is the final oxidation state of X?

A −1 **B** −2 **C** +2 **D** +3

6 (DHS 12) How many moles of electrons must be removed from each mole of methylbenzene, $C_6H_5CH_3$, when it is oxidised to benzoic acid, C_6H_5COOH?

A 1
B 2
C 4
D 6

7 (RV 12) An ion of metal **M** can be oxidised by potassium manganate(VII) in acid solution to form MO_3^-. In an experiment, 1.25×10^{-3} mol of the ion of **M** required 37.5 cm^3 of 0.0200 mol dm^{-3} potassium manganate(VII) for complete reaction.

What is the initial oxidation state of the ion of **M**?

A +1
B +2
C +3
D +4

8 (SAJC 12) Which species is oxidised and which species is reduced in the following reaction?

$$4TiCr_2O_4 + 8Rb_2CO_3 + 7O_2 \rightarrow 8Rb_2CrO_4 + 2Ti_2O_3 + 8CO_2$$

	Species oxidised	**Species reduced**
A	Chromium only	Carbonate ion only
B	Chromium only	Carbonate ion and oxygen
C	Chromium and titanium	Oxygen only
D	Chromium and titanium	Carbonate ion and oxygen

9 (TJC 12) 25.0 cm^3 of 0.0200 mol dm^{-3} of yellow **FA1** solution reacts with ***x*** g of zinc and the latter is oxidised to Zn^{2+}. The resulting solution required 15.0 cm^3 of 0.0200 mol dm^{-3} of acidified $KMnO_4$ to restore its original colour. Find ***x***.

A 0.0196 g
B 0.0218 g
C 0.0327 g
D 0.0491 g

10 (DHS 13) SO_2 gas is found in air near volcanoes. 218.4 cm^3 of SO_2 collected at s.t.p. near a volcano was reacted with 25.0 cm^3 of 0.130 mol dm^{-3} acidified potassium dichromate solution to turn the solution from orange to green. What is the oxidation state of sulfur at the end of the reaction?

A 0
B +2
C +4
D +6

11 (RV 13) In an experiment, 50.0 cm^3 of a 0.100 mol dm^{-3} solution of a metallic ion, M^{n+}, reacted exactly with 25.0 cm^3 of 0.100 mol dm^{-3} aqueous sodium sulfite.

The half-equation for the oxidation of sulfite ion is shown below.

$$SO_3^{2-}(aq) + H_2O(l) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

If the final oxidation number of the metal in the salt was +3, what would be the original oxidation number of the metal?

A +1
B +2
C +4
D +5

12 (SAJC 13) 0.0100 mol of an oxide of an element, **L**, represented by $\mathbf{L}_2O_n$ is found to react with exactly 8.00 x 10^{-3} mol of acidified $KMnO_4$ solution. In the reaction, aqueous $H\mathbf{L}O_4^{2-}$ ion is formed. What is the initial oxidation state of **L**?

A + 1
B + 2
C + 3

D + 4

13 (TJC 13) Comproportionation is a chemical reaction where two reactants, each containing the same element but with a different oxidation number, form a product in which the element involved reaches the same oxidation number in the product.
1 mol of potassium iodide undergoes comproportionation with compound **X** to form 0.6 mol of iodine. What could be the identity of compound **X**?

A KIO_2
B IO_2
C KIO_3
D KIO_4

14 (VJC 13) The equations for three reactions are given below.

$$Cl_2 + 2I^- \dashrightarrow 2Cl^- + I_2$$
$$2Fe^{3+} + 2I^- \dashrightarrow 2Fe^{2+} + I_2$$
$$Cl_2 + 2Fe^{2+} \dashrightarrow 2Cl^- + 2Fe^{3+}$$

What is the correct order of strength of Cl^-, Fe^{2+} and I^- as reducing agents?

	weakest	----->	*strongest*
A	Cl^-	I^-	Fe^{2+}
B	Cl^-	Fe^{2+}	I^-
C	Fe^{2+}	I^-	Cl^-
D	I^-	Cl^-	Fe^{2+}

15 (DHS 14) In an aqueous solution, 1 mol of MnO_4^- ions oxidises 3 mol of Fe^{2+} ions.
What is the manganese–containing product in this reaction?

A Mn^{2+}
B MnO_4^{2-}
C Mn^{3+}
D MnO_2

16 (NJC 13) When an excess of KI was added to 5.0 cm^3 of a 0.400 mol dm^{-3} solution of $\mathbf{A}O_4^{3-}$, 8.00 cm^3 of 0.500 mol dm^{-3} of $Na_2S_2O_3$ is required to react completely with the iodine liberated.
What is the final oxidation state of **A**?

A −3

B +1

C +2

D +3

17 (NYJC 14) Oxidation of atmospheric SO_2 can be catalysed by NO_2 as shown in the following equation.

$$SO_2(g) + 1/2O_2 \xrightarrow{NO_2(g)} SO_3(g)$$

The proposed mechanism for the oxidation process is as follows:

Step 1: $SO_2 + NO_2 \dashrightarrow SO_3 + NO$

Step 2: $NO + ½O_2 \dashrightarrow NO_2$

Which of the following statement is **not** correct?

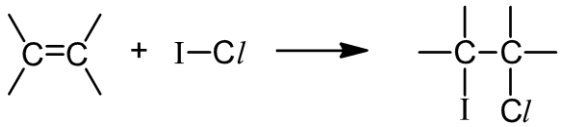
A Both SO_2 and SO_3 can cause acid rain.

B NO_2 is a free radical.

C NO is not a homogenous catalyst.

D Nitrogen atom exhibits variable oxidation states in this reaction.

18 (RI 14) The degree of unsaturation of an unknown fatty acid is usually determined by its iodine number, which is the mass of iodine atoms that add across all C=C bonds in 100 g of fatty acid in the following addition reaction. Iodine is introduced to the fatty acids in the form of IC*l* as IC*l* is more reactive than I_2.



In one reaction, 1.72 x 10^{-3} mol of IC*l* (in excess) were added to 0.100 g of fatty acid, and the excess IC*l* is converted to I_2 as follows:

$$ICl + KI \rightarrow KCl + I_2$$

The I_2 produced was titrated against sodium thiosulfate, $Na_2S_2O_3$.

$$I_2 + 2Na_2S_2O_3 \rightarrow 2NaI + Na_2S_4O_6$$

It was found that 1.39 x 10^{-3} mol of $Na_2S_2O_3$ were required for complete titration. What is the iodine number of the unknown fatty acid?

A 88

B 130

C 177

D 260

19 (SAJC 14) Which of the following are disproportionation reactions?

1 $3Cl_2 + 6NaOH \rightarrow 5NaCl + NaClO_3 + 3H_2O$

2 $IO_3^- + 5I^- + 6H^+ \rightarrow 3I_2 + 3H_2O$

3 $15Se + SeCl_4 + 4AlCl_3 \rightarrow 2Se_8[AlCl_4]_2$

A 1, 2 and 3
B 1 and 2 only
C 2 and 3 only
D 1 only

20 (TJC 14) In an experiment, 10 cm^3 of a solution containing a known concentration of iodate(V) ions, IO_3^-, was run into an acidified solution containing an excess of iodide ions.

$$IO_3^- (aq) + 5I^- (aq) + 6H^+ (aq) + 3H_2O (l) + 3I_2 (aq)$$

The iodine liberated was found on titration to react with exactly 30 cm^3 of 0.1 mol dm^{-3} of sodium thiosulfate solution.

What was the concentration of the iodate(V) ions in mol dm^{-3}?

A 0.05
B 0.10
C 0.15
D 0.20

21 (TJC 14) Aqueous acidified potassium manganate(VII) is used to oxidise iron(II) ethanedioate, FeC_2O_4, to Fe^{3+} and CO_2.

What is the concentration in mol dm^{-3} of 25 cm^3 aqueous acidified potassium manganate(VII) required to completely oxidise 0.05 mole of FeC_2O_4?

A 0.4
B 0.8
C 1.2
D 3.33

Structured Questions

1 (a) [TPJC/07]

When potassium nitrate, KNO_3, is heated, it decomposes according to the equation:

$$2\ KNO_{3(}s) \rightarrow\ 2\ KNO_2(s) + O_2(g)$$

Calculate the maximum mass of potassium nitrite, KNO_2, which could be obtained if 1.55 g of KNO_3 were *fully* decomposed by heat.

[1]

(b) Potassium nitrite, KNO_2, can be oxidised to potassium nitrate, KNO_3, by reacting it with acidified potassium manganate(VII).

(i) Construct a half-equation for the oxidation of NO_2^- to NO_3^-.

(ii) Hence, write a balanced equation for the reaction between KNO_2 and $KMnO_4$. [2]

(c) Another 1.55 g sample of KNO_3 was *partially* decomposed by heating. The residue was dissolved in water and the solution made up to 250 cm^3. A 25.0 cm^3 portion of this solution was found to react, in acidic medium, with 28.90 cm^3 of 0.0150 mol dm^{-3} $KMnO_4$.

(i) Calculate the mass of KNO_2 present in the residue.

(ii) Use the result obtained in parts (**a)** and **(c)(i)** to calculate the percentage conversion of KNO_3 to KNO_2 in this experiment.

[3]

2 *Use of the Data Booklet is relevant to this question.*

A 1.00 g sample of steel was dissolved in excess nitric acid. The resulting solution was found to contain manganese(II) ions. On addition of sodium bismuthate, $NaBiO_3$, all the manganese(II) ions were oxidised to manganate(VII) ions. After removing any excess bismuthate ions, the resulting purple solution was titrated with a solution containing iron(II) ions of concentration 0.10 mol dm^{-3}. 36.00 cm^3 of this solution was required to reach the endpoint, in which iron(II) ions were oxidised to iron(III) ions.

(a) What is the oxidation state of bismuth in $NaBiO_3$? [1]

(b) **(i)** Given that the bismuthate ion is reduced to Bi^{3+}, write the half-equation for this reduction reaction.

(ii) Hence, write an overall balanced equation for the oxidation of manganese(II) to manganate(VII) by bismuthate ions in acidic solution.

[2]

(c) Calculate the percentage by mass of manganese in the steel sample.

[3]

3 Household bleach contains sodium chlorate(I) , NaC*l*O. When bleach solution is acidified with dilute hydrochloric acid, chlorine is produced by the following reaction:

$$ClO^-(aq) + 2H^+(aq) + Cl^-(aq) \rightarrow Cl_2(aq) + H_2O(l)$$

Chlorine oxidises iodide to iodine, and the amount of iodine produced can be estimated using a standard sodium thiosulfate solution.

A 25.0 cm^3 sample of a household bleach is diluted to 250 cm^3. To a 25.0 cm^3 portion of this diluted solution, dilute hydrochloric acid is added. This is followed by the addition of excess potassium iodide solution. The resultant mixture is then titrated against 0.200 mol dm^{-3} sodium thiosulfate solution. The volume required is 18.50 cm^3.

Calculate the concentration of sodium chlorate(I) in the household bleach.

4 [SRJC/08]

Potassium manganate(VII) is an oxidising agent used in redox titrations.

A 50.0 cm^3 sample of iron(II) oxalate, FeC_2O_4, was dissolved in water and the solution made up to 250 cm^3. A 25.0 cm^3 portion of this solution was acidified and titrated with 20.80 cm^3 of 0.025 mol dm^{-3} potassium manganate(VII).

(i) State the change in oxidation number for manganese

(ii) Write an ionic equation for the oxidation of oxalate ion to carbon dioxide.

(iii) Write an overall equation for the reaction between potassium manganate(VII) and iron(II) oxalate.

(iv) Calculate the concentration of iron(II) oxalate in the original sample.

(v) State the colour change at the end point of the titration.

[Total:5]

5 [YJC/08]

An effective way of getting rid of any unpleasant taste and odour in water is the use of a synthetic yellowish green gas, chlorine dioxide, ClO_2. This gas is synthesised using sodium chlorate(V) and hydrogen peroxide in an acidic solution.

(a) Write a balanced ionic equation for the reaction between chlorate(V) ions and hydrogen peroxide.

[1]

(b) It has been found that when water becomes alkaline, chlorine dioxide disproportionates to form Cl^- and ClO_3^-. A precipitate, AgCl, is formed when 20 cm^3 of 0.05 mol dm^{-3} of $AgNO_3$(aq) is added.

$$6OH^- + 6ClO_2 \longrightarrow Cl^- + 5ClO_3^- + 3H_2O$$

$$Cl^- + Ag^+ \longrightarrow AgCl$$

(i) Explain what disproportionation means. [1]

(ii) Using the above information, calculate the volume of ClO_2 disproportionated at room temperature and pressure. [2]

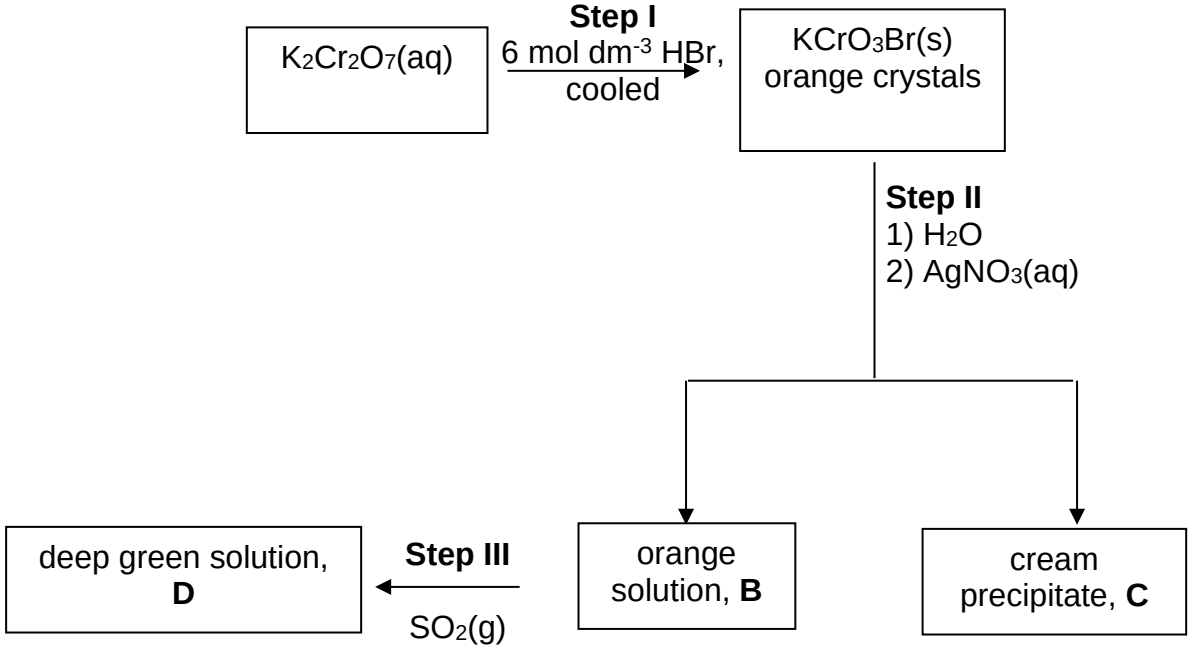
(iii) Write an overall equation for the reaction between potassium manganate(VII) and iron(II) oxalate.

(iv) Calculate the concentration of iron(II) oxalate in the original sample.

(v) State the colour change at the end point of the titration.

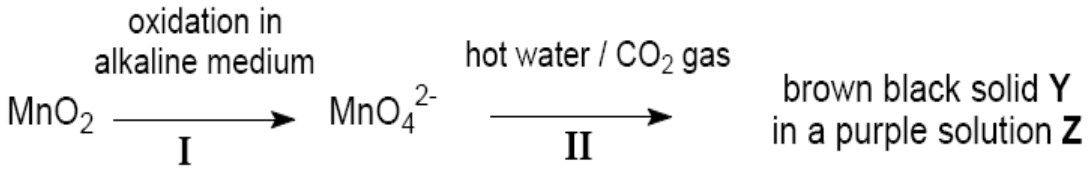
[5]

5 Potassium dichromate(VI) undergoes the following reactions.



(i) Write the equation for the reaction involved in **Step I**.
(ii) Is the reaction in **Step I** a redox one? Give your reasoning.
(iii) Identify compound **C**.

7 Manganese is commonly found in the ore pyrolusite, containing mainly manganese(IV) oxide. Manganese(IV) oxide undergoes a two-step reaction under alkaline condition to produce compounds **Y** and **Z**.



(i) The brown solid **Y** contains 63.8% by mass of manganese and 36.2% by mass oxygen. Using these information, determine the empirical formula of **Y**.

(ii) Suggest the identity of **Z** and state the type of reaction that occurs at step **II**. Construct a balanced equation for the reaction.

8 (NJC 12) The Dissolved Oxygen Concentration (DOC) in rivers and lakes is important for aquatic life. If DOC falls below 5 mg dm^{-3}, most species of fish cannot survive.

A sample of water is shaken with aqueous Mn^{2+} and aqueous alkali. The dissolved oxygen oxidises the Mn^{2+} to Mn^{3+}, forming a pale brown precipitate of $Mn(OH)_3$.

$$O_2(aq) + Mn^{2+}(aq) + 8\ OH^-(aq) + 2\ H_2O(l) \rightarrow 4\ Mn(OH)_3(s)$$

The $Mn(OH)_3$ precipitate is then reacted with an excess of aqueous potassium iodide, which is oxidised to iodine, I_2.

$$2\ Mn(OH)_3(s) + 2I^- \rightarrow I_2(aq) + 2\ Mn(OH)_2(s) + 2\ OH^-(aq)$$

The iodine formed is then determined by titration using aqueous sodium thiosulfate, $Na_2S_2O_3$ (aq).

A student decided to analyse the National Junior College's pond water using the above procedure. The student found that a 20.0 cm^3 sample of the pond water required 18.60 cm^3 of 0.00100 mol dm^{-3} of $Na_2S_2O_3$ (aq) for titration.

(i) Calculate the DOC of the sample of pond water, in mg dm^{-3}.

(ii) Comment on whether there is enough dissolved oxygen in the pond water for the carps to survive.

9 [RV 13]
The Claus process is used in the industry to partially oxidise H_2S to form water and elemental sulfur, S_8. Write a balanced equation and calculate the mass of elemental sulfur produced when 1 tonne of H_2S is used.
[1 tonne = 1000 kg]

Answers: 2) Redox
Redox

(1	**B**	**(2)**	**A**	**(7)**	**B**	**(10)**	**D**	**(13)**	**C**	**(16)**	**D**	**(19)**	**D**
(3	**C**	**(4)**	**A**	**(8)**	**C**	**(11)**	**C**	**(14)**	**B**	**(17)**	**C**	**(20)**	**A**
(5	**D**	**(6)**	**D**	**(9)**	**D**	**(12)**	**C**	**(15)**	**D**	**(18)**	**B**	**(21)**	**C**

3) Atomic Structure

MCQ

1 Which of the following contains only one unpaired electron?

A Ga^- **B** Se^- **C** Te^+ **D** As^+

2 Which atom has the highest ratio of unpaired electrons to paired electrons in its ground state?

A Boron

B Carbon

C Nitrogen

D Oxygen

3 Bombardment of ^{14}N with α particles (helium nuclei $24He$) produces a hydrogen nucleus and another nucleus. This other nucleus is

A $^{14}C_6$

B $^{14}N_7$

C $^{17}O_8$

D $^{18}Ne_{10}$

4 Chromium has four naturally occurring stable isotopes: ^{50}Cr, ^{52}Cr, ^{53}Cr, and ^{54}Cr.

Which statement about its ions is correct?

A $^{52}Cr^{3+}$ ion has less protons than $^{54}Cr^{3+}$ ion.

B The electronic configuration of $^{52}Cr^{3+}$ ion is $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^1$.

C The electronic configurations of $^{52}Cr^{3+}$ ion and $^{53}Cr^{3+}$ ion are different.

D $^{50}Cr^{3+}$ ion shows a greater angle of deflection than $^{52}Cr^{3+}$ ion in an electric field.

5 An element shows oxidation states of +4 and +6 in forming compounds with both fluorine and oxygen. What is the most likely outer electronic configuration for the element?

A $2s^2 2p^4$ **B** $2s^2 2p^6$ **C** $3s^2 3p^4$ **D** $3s^2 3p^6$

6 The successive ionisation energies (in kJ mol^{-1}) of an element are shown below:

762, 1540, 3300, 4390, 8950, 11900

Which Group does this element belong to?

A Group 13 **B** Group 14

C Group 15 **D** Group 16

7 **Figure 1** shows the first six ionisation energies of an element **X**.

Figure 2 shows the first ionisation energy of eight consecutive elements (including **X**) in Period 3.

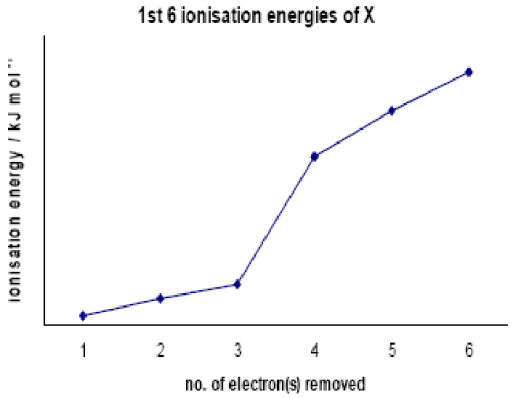


Figure 1

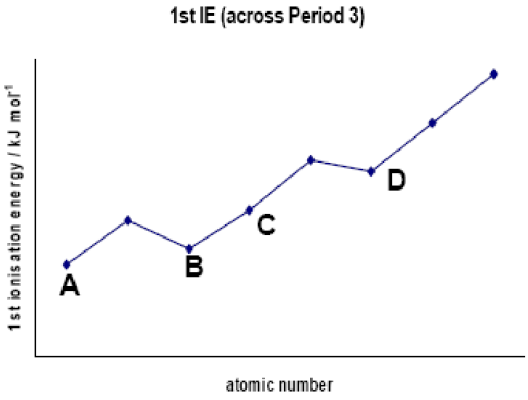
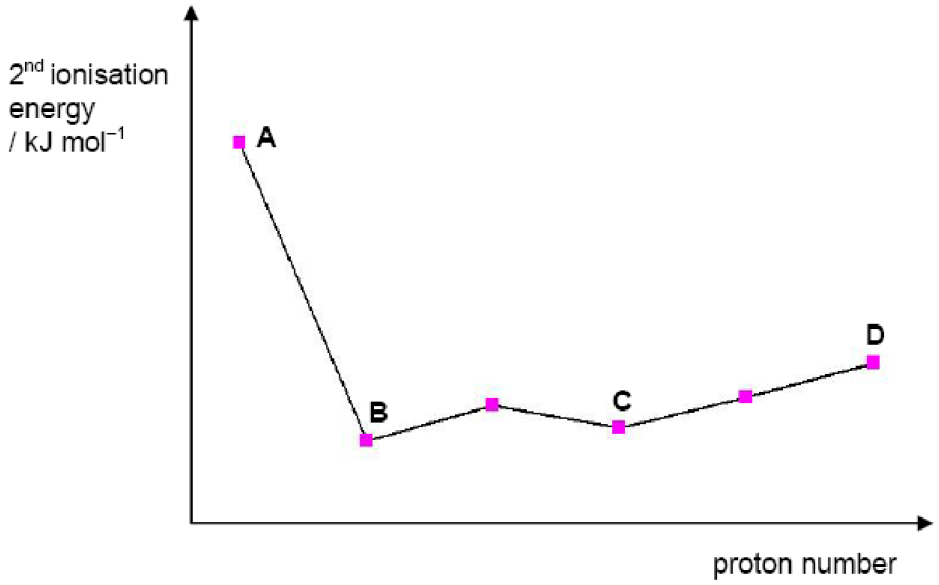


Figure 2

With reference to **Figure 2**, which option corresponds to the first ionisation energy of element **X**?

8 The variation in the second ionisation energy of six consecutive elements in the Periodic Table is shown in the graph.

Which of these elements is a Group 2 element?



9 Which statement best explains why sulfur has a lower 1st ionisation energy than phosphorus?

A The sulfur atom is larger in size than the phosphorus atom.

B Sulfur is more electronegative than phosphorus.

C The electron which is to be lost is paired with another electron.

D P has more protons than S so valence electrons in P experience greater nuclear attraction.

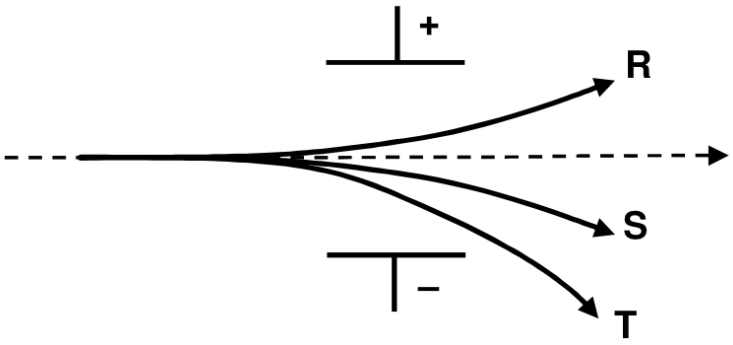
10 The successive ionisation energies, in kJ mol^{-1}, of an element **Z** are given below.

1010 2265 3662 4998 7002 8992 28698 32215

Which of the following would best represent the formula of a stable ionic compound formed by **Z**?

A **Z**F **B** C$\mathbf{Z}_2$ **C** Ca$\mathbf{Z}_2$ **D** Na_2**Z**

11 The following are flight paths of charged particles when accelerated in an electric field.



Which of the following are correct positions for the particles?

	R	S	T
A	$^{16}O^{+}$	$^{14}C^{+}$	$^{14}N^{+}$
B	$^{15}O^{-}$	$^{14}C^{+}$	$^{28}Si^{+}$
C	$^{14}N^{-}$	$^{15}O^{+}$	$^{28}Si^{2+}$
D	$^{14}N^{-}$	$^{14}C^{+}$	$^{28}Si^{4+}$

12 **A**, **B**, **C**, **D** and **E** are consecutive elements in the third period. The successive ionisation energies, in kJ mol^{-1}, of **B** are given below.

580 1820 2740 11600

Which statements are correct?

1 The 1st ionisation energy of **B** is higher than that of **A**.

2 $\mathbf{C^-}$ has a half–filled *p* subshell.

3 The compound formed from **D** and **E** is covalent.

A 1, 2 and 3

B 1 and 2 only

C 1 and 3 only

D 2 and 3 only

13 Which statements correctly describe an electron shell with the principal quantum number n = 2?

1 This shell can accommodate a maximum of eight electrons.

2 Each orbital can contain maximum two electrons with opposite spins.

3 An orbital in this shell may have spherical or dumb-bell shape.

A 1, 2 and 3

B 1 and 2 only

C 1 and 3 only

D 1 only

14 Which of the following particles have a single unpaired electron?

1 The copper ion in CuI

2 The methyl free radical

3 A molecule of NO

A 2 only

B 2 and 3 only

C 1 and 2 only

D 1, 2 and 3

15 Gaseous particle **X** has a proton number n and a charge of +1.

Gaseous particle **Y** has a proton number of (n + 1) and is isoelectronic with **X**.

Which of the following statements correctly describe **X** and **Y**?

1 **Y** has a smaller radius than **X**.

2 **Y** requires less energy than **X** when a further electron is removed from each particle.

3 **Y** releases less energy than **X** when an electron is added to each particle.

A 1, 2 and 3

B 1 and 2 only

C 1 and 3 only

D 1 only

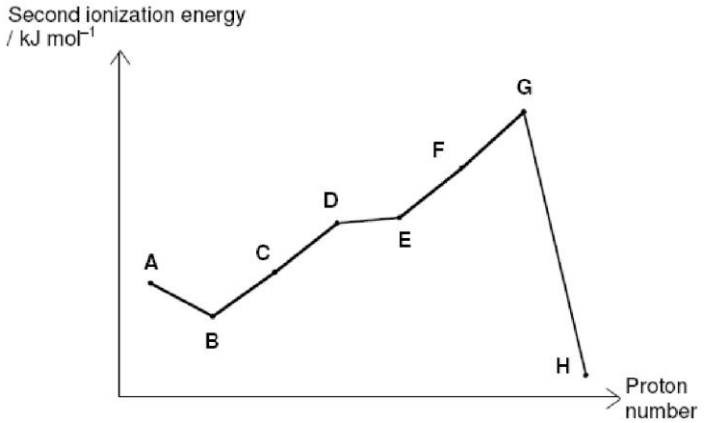
Structured Questions

1 Sulfur exists in its pure elemental state in nature as a yellow crystalline solid.

(a) Give the equation that represents the second ionisation energy of sulfur.

[1]

(b) The graph below shows the second ionisation energies of eight consecutive elements in the Periodic Table.

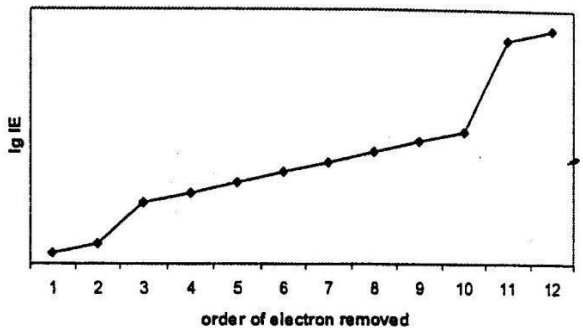


(i) Which of the above elements, **A** to **H**, is sulfur? Explain your answer.

(ii) Explain the trend of the second ionisation energies from **A** to **G**, including any irregularities.

[4] [HCI/07]

2 The diagram below shows successive ionisation energy values for an unknown element **Z**.



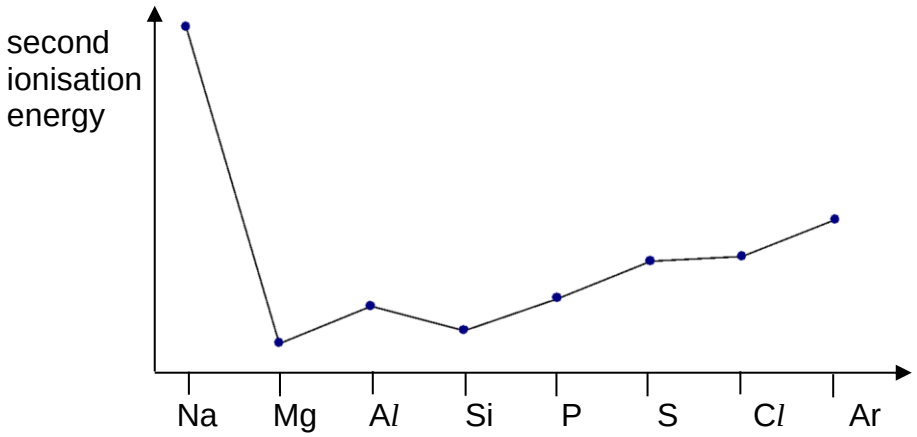
(i) Account for the large difference between the 2nd and 3rd ionisation energy values.

(ii) How many occupied quantum shells does **Z** have?

(iii) Write the electronic configuration of **Z.**

[3] [SRJC/07]

3(a) The **second ionisation energies** of the elements in Period 3 are given in the graph below.



(i) Define, by means of an equation, what is meant by the **second ionisation energy** of magnesium.

(ii) Explain why the second ionisation energy of silicon is lower than that of aluminium.

[2]

(b) Orange street lamps contain sodium metal with a small amount of neon gas. The light is produced when gaseous atoms are ionised in an electric field. When first turned on, the lamps emit a red glow characteristic of neon but, after a time, the orange glow of sodium predominates.

Suggest a brief explanation for the observations. [2]

[CJC09]

4(a) According to Quantum Chemistry, electrons occupy specific probability regions called orbitals. Draw and label the shapes of the two types of orbital found in oxygen atoms.

[2]

(b) State the electronic configuration of the

(i) manganese(II) ion, Mn^{2+} **(ii)** germanium atom, Ge.

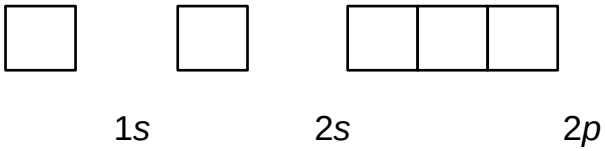
[2]

(c) Na, Mg and Al are three consecutive elements in Period 3. Explain why the first ionisation energy of Mg is higher than those of Na and Al.

[2]

5 Nitrogen, N_2 and carbon monoxide, CO are isoelectronic.

(i) The orbitals of a nitrogen atom may be represented as shown. Draw the arrangement of electrons in the ground state of a nitrogen atom.



(ii) Draw and label the shapes of all the orbitals found in a nitrogen atom.

Answers: 3) Atomic Structure

MCQ

1	B	5	C	9	C	13	A
2	C	6	B	10	D	14	B
3	C	7	B	11	D	15	D
4	D	8	B	12	D		

4) Physical Periodicity MCQ

1 Which equation corresponds to the third ionisation energy of titanium (Ti)?

A $Ti(g) \rightarrow Ti^{3+}(g) + 3e^-$

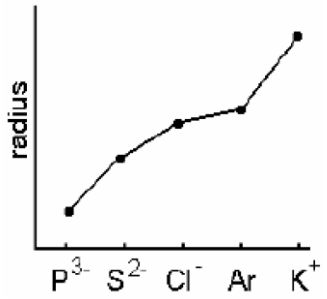
B $Ti^{2+}(s) \rightarrow Ti^{3+}(g) + e^-$

C $Ti^{2+}(g) \rightarrow Ti^{3+}(g) + e^-$

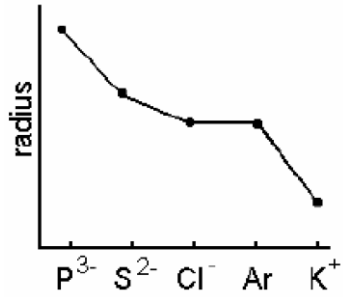
D $Ti^{3+}(g) + e^- \rightarrow Ti^{2+}(g)$

2 Which graph correctly shows the variation of the radius of the species shown?

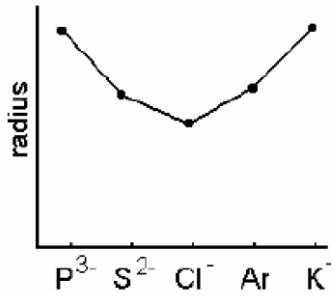
A



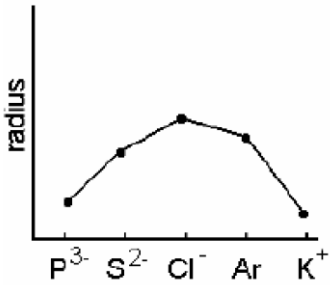
B



C



D



3 Which statement shows an increasing trend from left to right across the elements of the third period of the Periodic Table?

A The ionic radii of the ions.

B Melting points of the elements.

C First ionisation energy of the elements.

D Electronegativity of the atoms.

4 The successive ionisation energies (kJ mol^{-1}) of two elements are given below:

X	1012	1907	2914	4964	6274	21267	26431	29876	35905

Y	1251	2298	3822	5159	6542	9362	11108	33604	38860

Which of the following statements is **false**?

A The formula of the compound formed between X and Y is XY_3.

B The compound formed between X and Y is covalent.

C X has a giant covalent structure.

D Y exists as diatomic molecules.

5 Elements **W**, **X**, **Y** and **Z** are four consecutive elements in the Periodic Table. The following table shows the first four ionisation energies of each element.

Element	**1st I.E**	**2nd I.E**	**3rd I.E**	**4th I.E**
W	1685	3374	6050	8408
X	2081	3952	6122	9371
Y	496	4562	6910	9543
Z	738	1451	7733	10543

Which Groups do elements **W** and **Z** belong to?

	W	**Z**
A	Group 17	Group 2
B	Group 5	Group 1
C	Group 13	Group 16
D	Group 2	Group 14

6 Which of the following statements is correct?

A The ionic radius of Na^+ ion is greater than that of Cl^- ion.

B The electronegativity of phosphorus is greater than that of nitrogen.

C The first ionisation energy of magnesium is less than that of aluminium.

D The energy required to remove an electron from F^- is less than that for Ne atom.

7 The first six ionisation energies (in kJ mol^{-1}) of a Period 2 element, **Z**, are shown.

1090	2350	4610	6220	37831	47277

Which diagram represents the arrangement of electrons in the ion, $\mathbf{Z}^{2-}$?

	1s	2s	2p		
A	↿⇂	↿⇂	↿	↿	
B	↿⇂	↿	↿	↿	↿
C	↿⇂	↿⇂	↿⇂	↿	↿
D	↿⇂	↿⇂	↿⇂	↿⇂	↿⇂

8 In which one of the following pairs does the first species have a higher first ionisation energy than the second one?

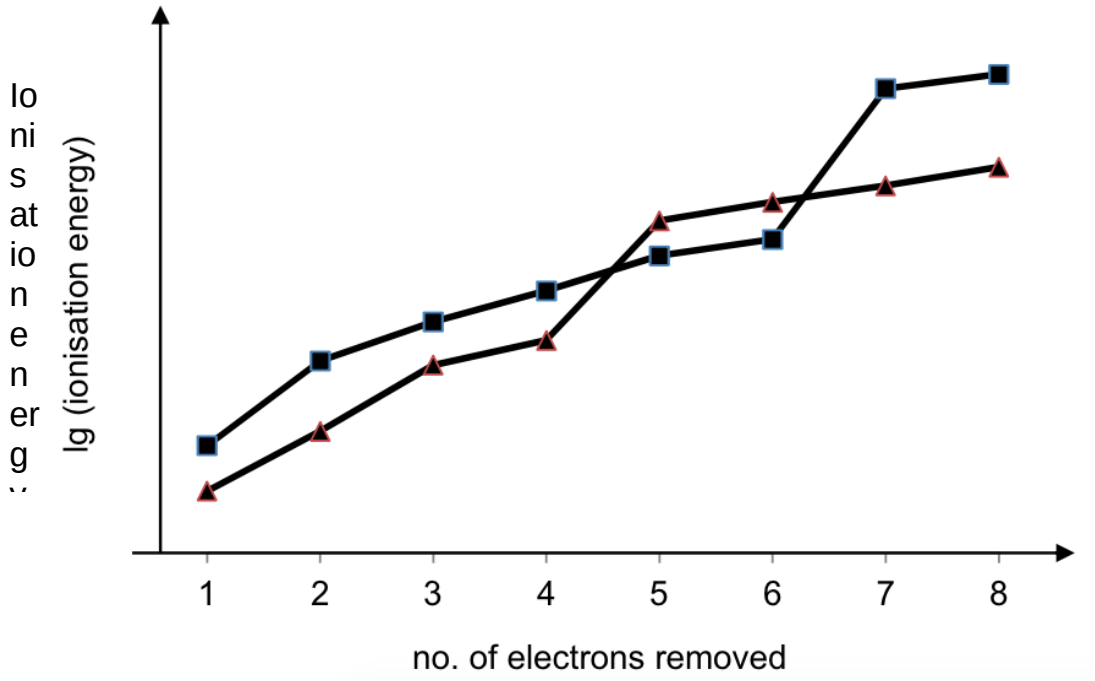
A Sb, Te

B Kr, Ar

C Mg^+, Al^+

D Se, Se^+

9 The graph shows the first eight ionisation energies of two elements in Periods 2 and 3.



What is the most likely compound that will be formed between the two elements?

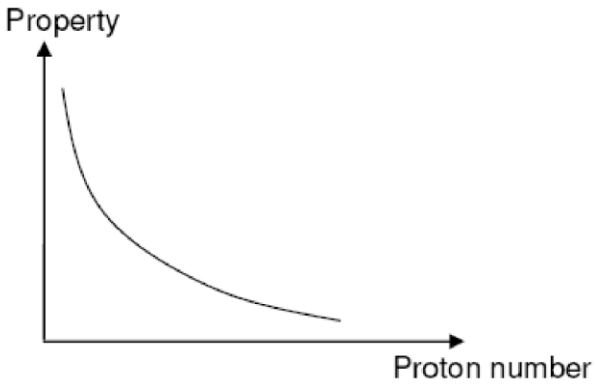
A CO_2

B SiO_2

C CS_2

D $SiCl_4$

10 The graph shows how a property of some elements varies with proton number.



Which of the following can be the property that shows the above trend?

1 Atomic radius of Period 2 elements

2 Melting point of Group 1 elements

3 Electronegativity of Group 17 elements

A 1, 2 and 3

B 1 and 2 only

C 1 and 3 only

D 1 only

11 *Use of the Data Booklet is relevant to this question.*

M and **N** have the following features:

- **M** has fewer unpaired electrons than **N**.
- **M** is smaller than **N**.

Which of the following pairs of **M** and **N** possess these features?

	M	**N**
1	S^-	P^-
2	Mg	A*l*
3	Zn^{2+}	Cu^+

A 1, 2 and 3

B 1 and 2 only

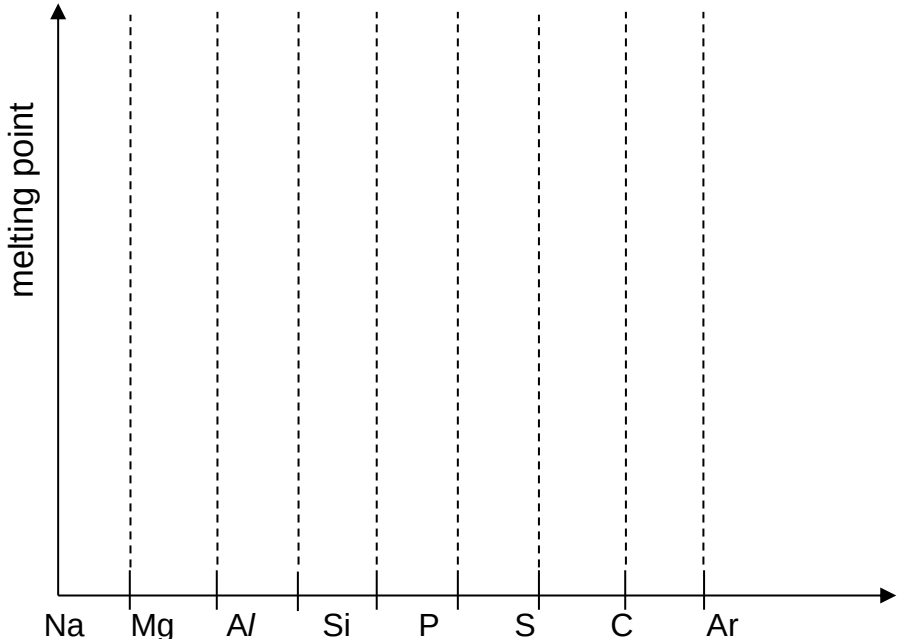
C 1 and 3 only

D 1 only

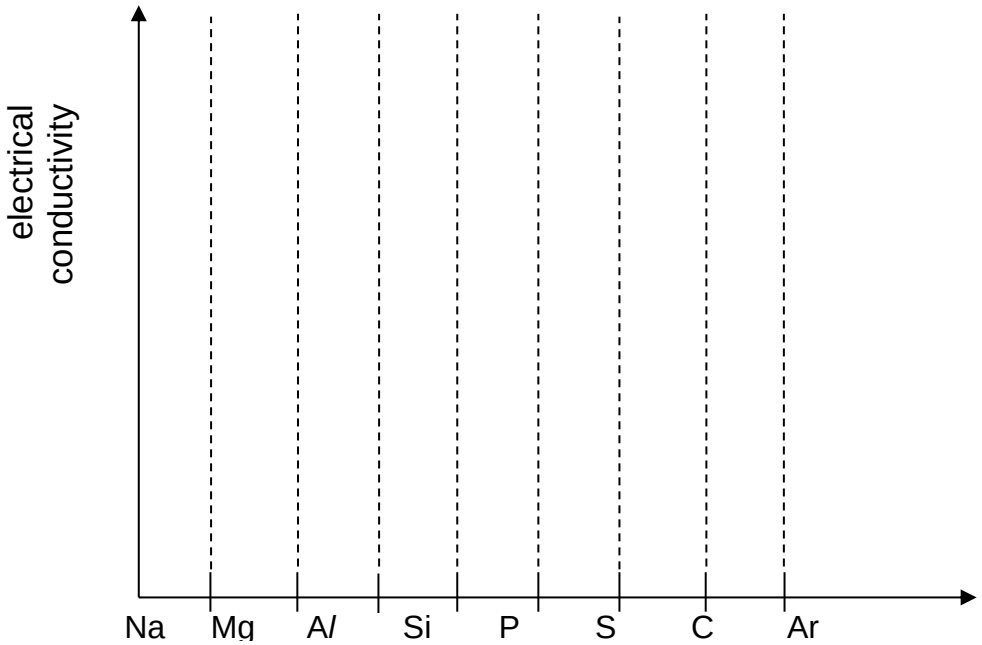
Structured Questions

1(a) On each of the grids below, sketch the trends of the properties of the elements across the third period of the Periodic Table. Below each grid, explain briefly why these properties vary in the way you have shown.

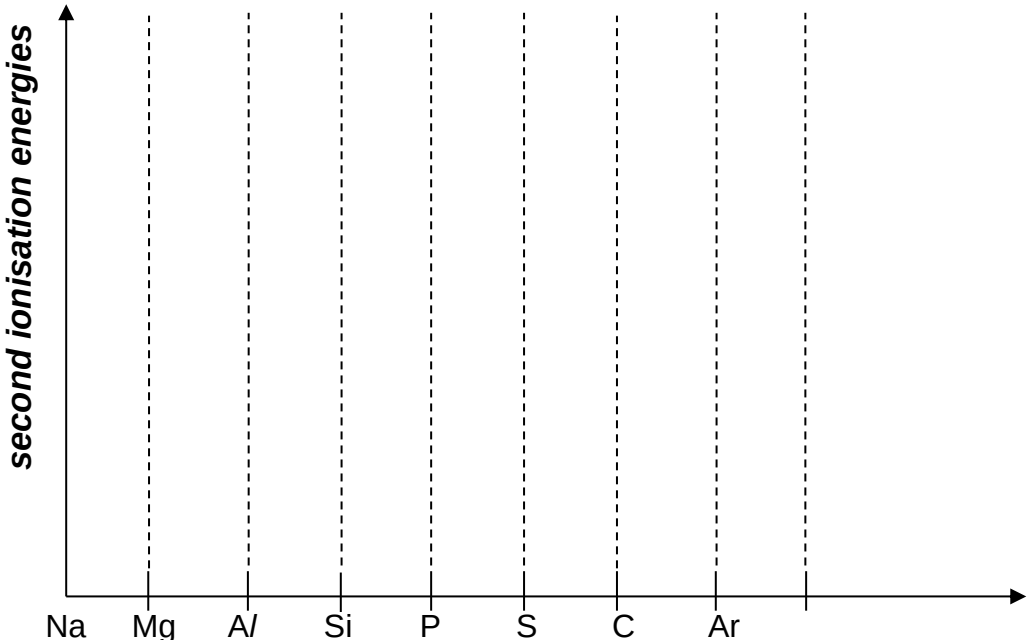
(i) melting point



(ii) electrical conductivity



(b)(i) In the grid below, sketch the variation of the ***second ionisation energies*** of the elements across the period.



(ii) Explain the difference between the second ionisation energies of aluminium and silicon.

[3][RJC/07]

2 The first, second and third ionisation energies of manganese and iron are given in the table below:

	First Ionisation Energy / kJ mol^{-1}	Second Ionisation Energy/ kJ mol^{-1}	Third Ionisation Energy/ kJ mol^{-1}
Manganese	716	1510	3250
Iron	762	1560	2960

(i) Why is there an increasing trend in successive ionisation energies?

(ii) State the electronic configurations of Mn^{2+} and Fe^{2+}.

(iii) Hence, suggest for the difference between the third ionisation energies of manganese and iron.

Answers: 4) Physical Periodicity

MCQ

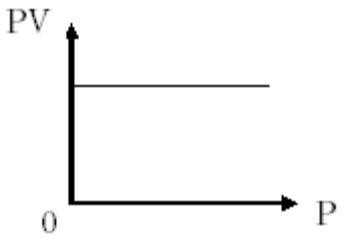
1	**C**	4	**C**	7	**C**	10	**A**
2	**B**	5	**A**	8	**A**	11	**D**
3	**D**	6	**D**	9	**B**		

5) Ideal Gases

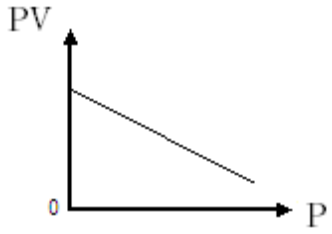
MCQ

1. A fixed mass of an ideal gas occupies a vessel of volume V at a pressure P. Which graph correctly shows the relationship between product PV and P? (Assuming gas is at constant temperature).

A

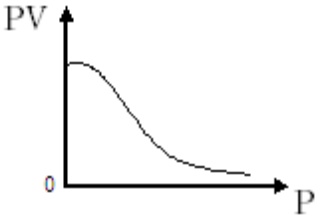


B



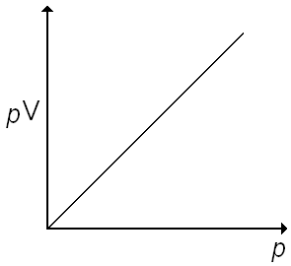
C

D

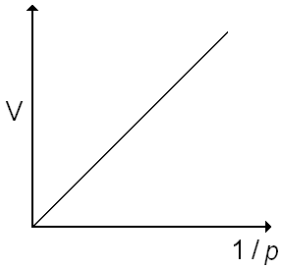


2. Which graph does **not** describe the relationship between volume and pressure for a fixed mass of an ideal gas under constant temperature?

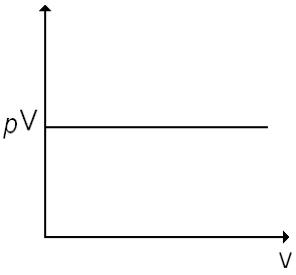
A



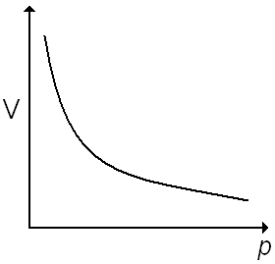
B



C

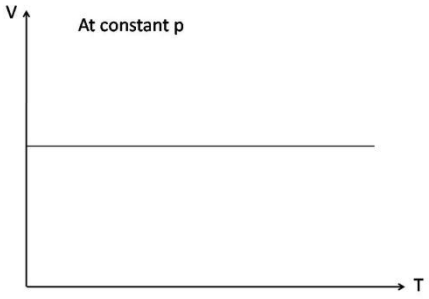


D

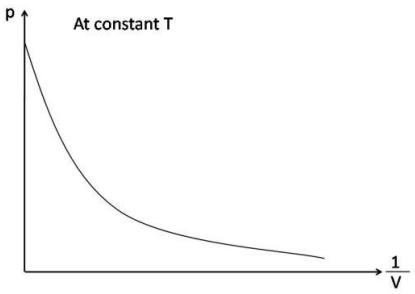


3. Which graph correctly describes the behavior of a fixed mass of an ideal gas?(T is measured in K.)

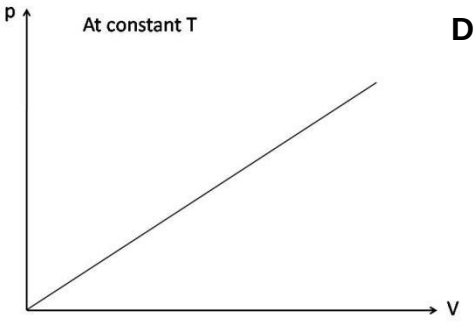
A



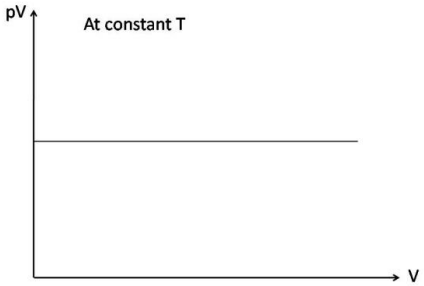
C



B

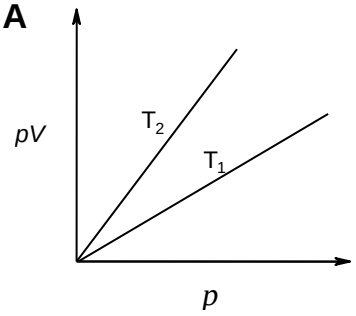


D

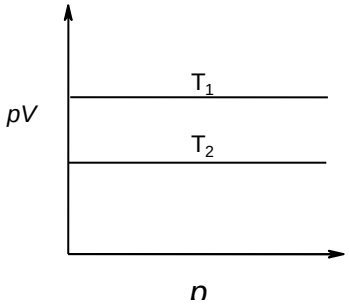


4. Which of the following graphs below shows the correct plot of pV against p for a fixed mass of an ideal gas at two temperatures, T_1 and T_2, where $T_1 > T_2$?

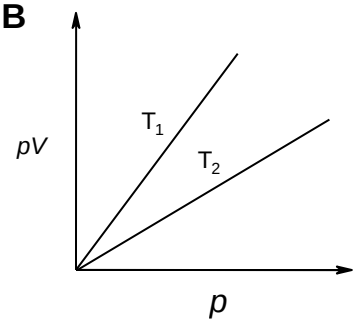
A



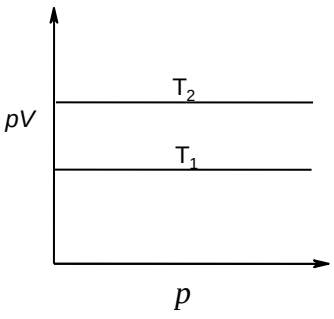
C



B

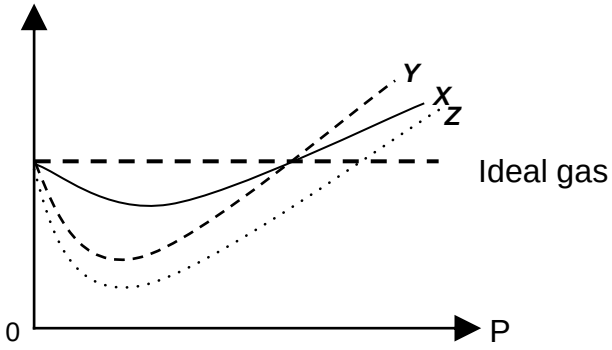


D



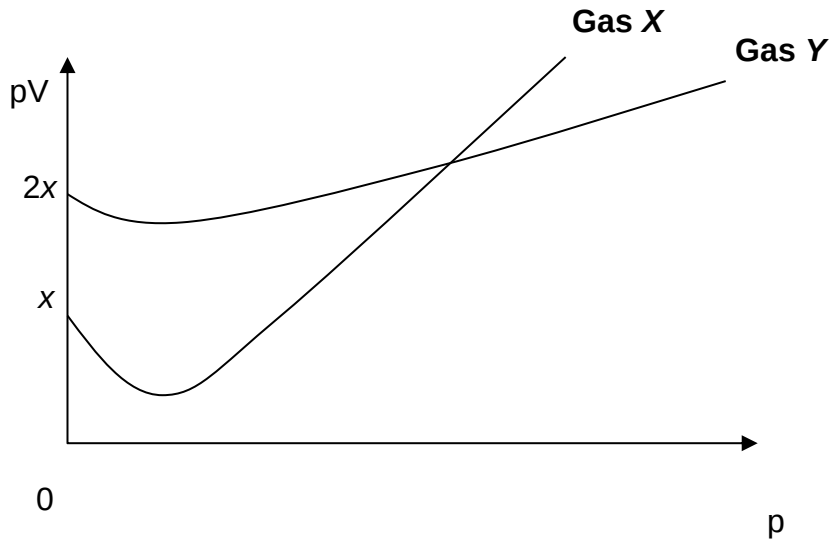
5. The following graph shows the comparison between ideal and real gases at constant temperature.

What are the identities of gases **X**, **Y** and **Z**?



	X	**Y**	**Z**
A	CO_2	NH_3	N_2
B	N_2	CO_2	NH_3
C	NH_3	N_2	CO_2
D	NH_3	CO_2	N_2

6. The value of pV is plotted against p for 2 gases, ***A*** and ***B***, where p is the pressure and V is the volume of the gas.



Which of the following could be the identities of the gases?

	Gas X	**Gas Y**
A	0.5 mol of H_2 at 25 ºC	0.5 mol of H_2 at 50 ºC
B	0.5 mol of H_2 at 25 ºC	1 mol of SO_2 at 25 ºC
C	0.25 mol of SO_2 at 25 ºC	0.5 mol of H_2 at 25 ºC
D	1 mol of SO_2 at 25 ºC	0.5 mol of H_2 at 25 ºC

7. Flask **X** contains 1 dm^3 of helium at a pressure of 2 kPa and flask **Y** contains 2 dm^3 of neon at a pressure of 1 kPa.

If the flasks are connected at constant temperature, what is the final pressure?

A $\frac{4}{3}$ kPa **B** $\frac{5}{3}$ kPa **C** $\frac{3}{2}$ kPa **D** 2 kPa

8. Three vessels containing 90 cm^3 of iodine fumes, 4 dm^3 of argon and 0.01 m^3 of water vapour respectively at 150 °C and 101 kPa were connected together. The connected vessels were allowed to cool to room temperature.

What was the final pressure in the connected vessel?

A 19.8 kPa **B** 71.1 kPa **C** 28.7 kPa **D** 70.7 kPa

9. A sample of gas in a flask contains 2 g of helium and 32 g of oxygen at 300 Pa. What is the partial pressure of helium in the flask?

A 60 Pa **B** 100 Pa **C** 150 Pa **D** 200 Pa

10. A flask **L** containing an unknown volume of neon gas at a pressure of 2 kPa was connected to a flask **M** containing 3 dm^3 of helium gas at 4 kPa. The final pressure in the connected flasks was found to be 3.5 kPa.

What was the volume of flask **L**, assuming that the temperature was constant throughout?

A 1.0 dm^3 **B** 1.5 dm^3 **C** 2.0 dm^3 **D** 2.5 dm^3

11. Under extreme conditions, gallium chloride Ga_2Cl_6 can be formed via the endothermic dimerisation of $GaCl_3$.

$$2GaCl_3(g) \rightleftharpoons Ga_2Cl_6(g)$$

In an experiment, a total pressure of 7.75 x 10^3 Nm^{-2} was measured when a 3.745 x 10^{-3} mol sample of $GaCl_3$ was allowed to dimerise in a 1.80 dm^3 vessel at 200 °C.

What is the fraction of $GaCl_3$ dimerised at 200 °C, assuming ideal gas behaviour?

A 0.053 **B** 0.106 **C** 0.894 **D** 0.947

12. A sample of 0.16 g of liquid **X** is vapourised in a graduated syringe at 100 °C and 1 atm. The initial volume of air in the syringe recorded is 10 cm^3 and the final recorded volume is 56 cm^3. What is the relative molecular mass of **X**?

A 23.4 **B** 87.4 **C** 88.6 **D** 106.4

13. A vessel contains a mixture of nitrogen and oxygen at 30 °C and a pressure of 200 kPa. After all oxygen has reacted with zinc, the total pressure in the vessel is 150 kPa whereas the temperature is 30 °C.

What is the molar ratio of nitrogen to oxygen in the mixture?

A 1:3 **B** 2:3 **C** 3:1 **D** 3:2

14. A certain mass of an ideal gas occupies a volume of 1 dm^3 at 20 ^oC.
If the pressure remains unchanged, at what temperature will its volume be 2 dm^3?

A 40 ^oC **B** 293 ^oC **C** 313 ^oC **D** 586 ^oC

15. A 2 g sample of hydrogen at temperature *T* and of volume *V* exerts a pressure *p*.

Deuterium, $^{2}_{1}H$, is an isotope of hydrogen.

Which of the following would also exert a pressure of *p* at the same temperature *T*?

A a mixture of 1 g of hydrogen and 1 g of deuterium of total volume *V*

B a mixture of 1 g of hydrogen and 2 g of deuterium of total volume *V*

C a mixture of 1 g of hydrogen and 2 g of deuterium of total volume 2*V*

D a mixture of 2 g of hydrogen and 2 g of deuterium of total volume 2*V*

16. Deuterium (^2H) and Tritium (^3H) are isotopes of hydrogen.
Given that a 3 g sample of hydrogen at temperature **T** and volume **V** exerts a pressure **p**, which of the following would also exert a pressure **p** at the same temperature **T**?

A 3 g of deuterium of volume **V**

B 5 g of tritium of volume $\frac{1}{3}$**V**

C A mixture of 1g of deuterium and 1g of tritium of total volume **V**

D A mixture of 4g of deuterium and 3g of tritium of total volume **V**

17. A sample of an organic compound of molar mass *M* is vapourised in a gas syringe and occupies *V* cm^3 at *T* K and *p* bar.
What is the density of the organic compound?

A $\rho = \frac{p \times M \times (T + 273)}{22700 \times 273}$

B $\rho = \frac{p \times M \times 273}{22700 \times T}$

C $\rho = \frac{p \times M}{22700 \times (273 + T)}$

D $\rho = \frac{M \times (T + 273)}{p \times 22700 \times 273}$

18. The volume of a balloon at the Earth's surface is 100 cm^3.
What is the volume of the balloon as it increases to an altitude where the temperature is 6 °C and pressure is 25 kPa? [You may assume r.t.p at the Earth's surface.]

A 0.362 cm^3 **C** 362 cm^3

B 97.0 cm^3 **D** 378 cm^3

19. A 14.5 mg of an unknown sample of boron hydride, BxHy was found to occupy a volume of 188 cm^3 at a pressure of 3.59 kPa and a temperature of 25 ^oC.
What is the molecular formula of the boron hydride?

A B_2H_6 **B** B_4H_{10} **C** B_5H_9 **D** B_6H_{10}

20. At 1.5 atm and 27 ^oC, 1 mole of gas **A** occupies a volume of 16.7 dm^3 while 1 mole of gas **B** occupies a volume of 10 dm^3.
Which statement best explains the observation?

A The gas molecules for gases **A** and **B** experience intermolecular van der Waals' forces of attraction.

B Gases **A** and **B** are ideal.

C Gas **A** may be He while gas **B** may be CO_2.

D Gas **A** may be NH_3 while gas **B** may be H_2.

21. Which statement best explains why gases are easily compressed?

A Gas molecules have negligible attractive forces for one another.

B The volume occupied by the gas is much greater than that occupied by the molecules.

C The average energy of the molecules in a gas is proportional to the absolute temperature of the gas.

D The collisions between gas molecules are elastic.

22. Under which set of conditions does ammonia gas deviate least from ideality?

A high temperature, high pressure **B** low temperature, high pressure

C high temperature, low pressure **D** low temperature, low pressure

23. At standard temperature and pressure, easily liquefiable gases show greater deviation from ideal behaviour compared to gases that are difficult to liquefy.
Which statement best explains this phenomenon?

A Larger molecular volume of easily liquefiable gases causes greater impacts on vessel walls during collisions. This leads to greater deviation from ideality.

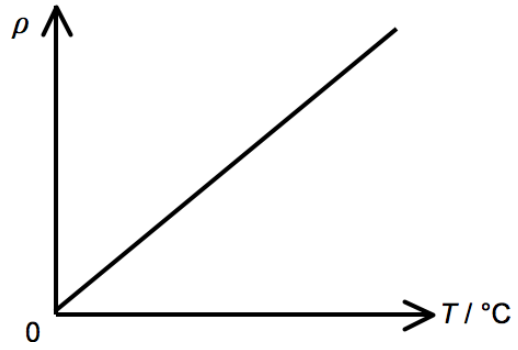
B For real gases, $pV > nRT$ due to repulsion between molecules. This effect is more significant for gases that are difficult to liquefy.

C The spread in kinetic energies is smaller in molecules of easily liquefiable gases, giving rise to greater deviation.

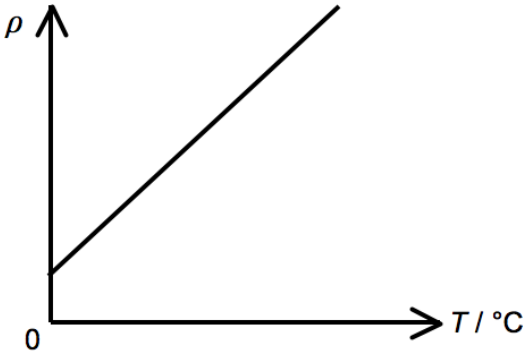
D Negligible intermolecular forces operate among molecules. This assumption is more valid for molecules that are difficult to liquefy

24. Which of the following graph is obtained when density (ρ) is plotted against temperature (T) for an ideal gas under constant pressure?

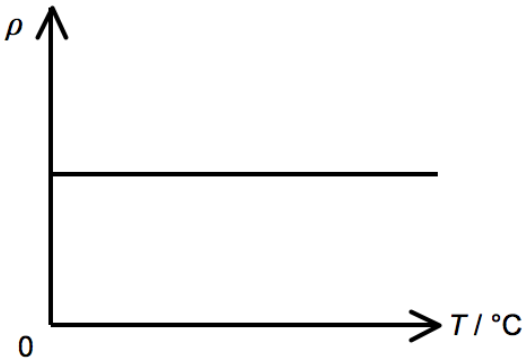
A



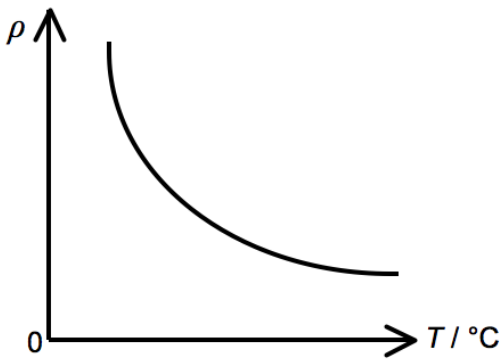
B



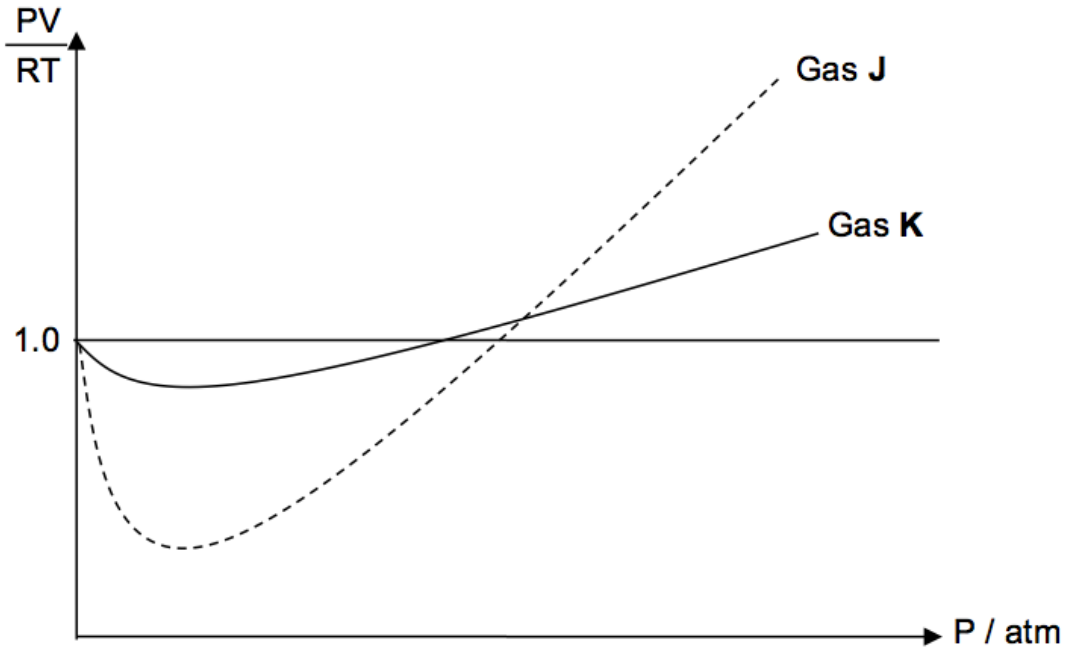
C



D



25. The graph below shows the variation in the values of PV/RT with pressure for 1 mol of Gas **J** and 1 mol of Gas **K**.



Which of the following is true of the identity of Gas **J** and Gas **K**?

	Gas **J**	Gas **K**
A	Neon at 498 K	Neon at 298 K
B	Hydrogen gas at 298 K	Carbon dioxide at 298 K
C	Neon gas at 498K	Carbon dioxide at 498 K
D	Nitrogen gas at 298K	Nitrogen gas at 498K

26. Which of the following gases shows the greatest deviation from ideality?

A NH_3

B CH_4

C C_2H_4

D N_2

28. The Gas Laws can be summarised with the ideal gas equation: pV = nRT

Which of the following statements are correct?

1 The graph of pV/T against V is a horizontal line for a given mass of an ideal gas.

2 The density of a given mass of an ideal gas at constant pressure is inversely proportional to the temperature.

3 The volume of a given mass of an ideal gas is doubled if its temperature is raised from 25 °C to 50 °C at constant pressure.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

29. In the ideal gas equation of *PV* = *nRT,* each symbol has its usual meaning.

Which of the following statements are correct?

1 Real gas deviates most from ideal gas behaviour at high pressure and low temperature.

2 The density of an ideal gas at constant pressure is inversely proportional to the absolute temperature.

3 The volume of the gas of an ideal gas is doubled when its temperature is raised from 25 ^{o}C to 50 ^{o}C at constant pressure.

A 1, 2 and 3
B 1 and 2 only
C 2 and 3 only
D 1 only

30. Which of the following statements are about a gas is believed to behave ideally and is described exactly by the equation pV = nRT?

1 The gas cannot be liquefied.

2 The pressure of a fixed mass of the gas is inversely proportional to its volume at constant temperature.

3 There are intermolecular forces of attraction in the gas.

A 1, 2 and 3
B 1 and 2 only
C 2 and 3 only
D 1 only

31. **V** and **W** are ideal gases that do not react together. The mass of 1 mol of **V** is four times that of **W**.

Which statements are true about **V** and **W** at s.t.p?

1 On mixing 1 dm^3 of **V** with 1 dm^3 of **W**, the partial pressure of each gas in the mixture will be approximately 50 kPa.

2 The mass of 1 dm^3 of **V** is four times that of 1 dm^3 of **W**.

3 The average kinetic energy of gas molecules of **V** is equal to that of **W**.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

32. A container of volume 30 dm^3 contains ideal gas **A** at a pressure of 5 atm. Another container of volume 20 dm^3 contains ideal gas **B** at a pressure of 15 atm. When the containers are connected at constant temperature, both gases **A** and **B** flow freely throughout both containers.

Which of the statements correctly describe the gases in the combined containers?

1 The partial pressure of gas **A** is 3 atm if the gases do not react.

2 The total pressure after connection is less than 9 atm if the gases react according to the equation **A(g) + B(g)** □ **C(g)**.

3 The total pressure in the containers after the connection is 20 atm if the gases do not react.

A 1, 2 and 3 only

B 1 and 2 only

C 2 and 3 only

D 1 only

33. P_{He} is the partial pressure of helium and P_{Ne} is the partial pressure of neon.

P is the total pressure of the mixture of the two gases.

Which of the following are correct about a mixture containing 20 g of helium and 20 g of neon? [He, 4; Ne, 20]

1 $P_{He} = 0.5P$ **2** $P_{He} + P_{Ne} = P$ **3** $P_{He} = 5P_{Ne}$

A 1, 2 and 3 only

B 1 and 2 only

C 2 and 3 only

D 1 only

34. Which of the following explains why real gas deviate from ideal gas behaviour?

1 Gas particles are simple covalent molecules with weak intermolecular van der Waals' forces of attraction.

2 Gas particles can be compressed to a certain minimum volume, beyond which cannot decrease further even if pressure increases.

3 Gas particles move at higher speeds when temperature increases.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

Structured Questions

1(a) 0.200 mol of radon is stored in a 20.0 dm^3 flask with 0.800 mol of fluorine at 127 ^oC.

(i) Calculate the total pressure in the flask.

(ii) Calculate the partial pressure of radon and fluorine in the flask.

(iii) Radon and xenon are in the same group. A scientist hypothesised that radon has similar chemical reaction as xenon. Xenon can react with fluorine according to the following reaction.

$$Xe(g) + F_2(g) \rightarrow XeF_2(g)$$

The mixture of radon and fluorine in the flask was sparked at 127 ^oC. Calculate the final pressure in the flask?

2(a) The engine capacity of a car refers to the maximum total volume of all the combustion cylinders when the pistons are pushed out. Normally, a fuel-air mixture of a specific ratio is injected into a cylinder and compressed to a small volume before being ignited. The combustion products expand and push out the piston to the maximum volume.

An experimental four-cylinder engine where each cylinder has a maximum volume of 400 cm^3 was used for the following process:

Vapourised fuel with formula C_8H_{18} was mixed with air and injected into each cylinder. The resultant mixture was compressed to a volume of 115 cm^3 at a pressure of 12.5 atm and temperature of 298 K.

The fuel-air mixture was ignited and the resultant combustion products caused the piston to be pushed out to bring the volume of the cylinder to the maximum, with a final pressure of 12.5 atm at 986 K.

(i) Calculate the resultant volume when 115 cm^3 of the gas mixture is heated from 298 K to 986 K at constant pressure of 12.5 atm. (1 atm = 101.3 kPa)

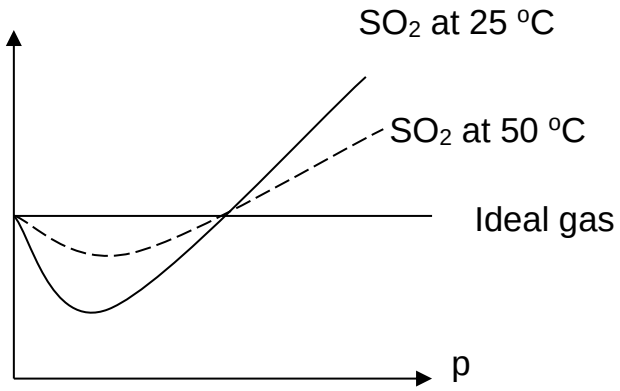
(ii) Using your result in **(a)(i)**, calculate the number of moles of gas at 986 K just before ignition.

(iii) Hence, determine the increase in the number of moles of gaseous species in each cylinder after combustion.

(iv) Write the equation, including state symbols, for the complete combustion of C_8H_{18} at 986 K and hence determine the number of moles of C_8H_{18} combusted.

3(a) With the help of the *Data Booklet*, determine the density, in g cm^{-3}, of $SO_2(g)$ at 25 ^oC and 1 atm.

(b) The following graphs show the behaviour of the same amount of $SO_2(g)$ at various temperatures. Explain the behaviour of the gas with increasing pressure at the higher temperature.



4. Hydrazine is a chemical compound with the formula N_2H_4. It has an ammonia-like odour, and is derived from the same industrial chemistry processes that

manufacture ammonia. However, hydrazine has physical properties that are more similar to those of water. Hydrazine is used as rocket fuel and to prepare the gas precursors used in air bags.

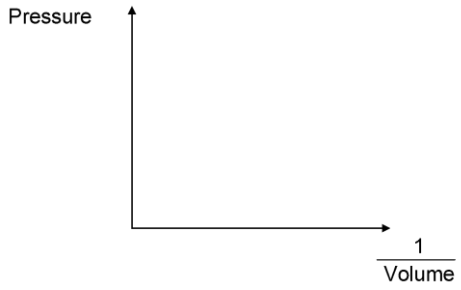
(a) Draw the Lewis structure for the hydrazine molecule. State its shape and indicate the H–N–H bond angle.

(b)(i) Calculate the volume occupied by 1 mole of ideal gas at a pressure of 1.5 atm and 300 K.

(ii) At a pressure of 1.5 atm and 300 K, 1 mole of $N_2H_4(g)$ occupies 15.0 dm^3. Explain the difference in the volume occupied by hydrazine and the volume calculated in **(i)**.

(iii) Under what conditions of temperature and pressure would you expect $N_2H_4(g)$ to occupy the same volume as that calculated in **(i)**?

(iv) Assuming that $N_2H_4(g)$ is placed under the conditions specified in **(b)**(**iii)**, sketch the graph of pressure against the reciprocal of volume for a given amount of hydrazine at constant temperature.



5. One of the uses of carbon dioxide is in the manufacture of carbonated drinks.

Cola, a carbonated drink, was originally invented by chemist, John Pemberton, to stop headaches and calm nervousness.

Cylinders of pressurised carbon dioxide are used to produce cola. A commercial cola drink was manufactured using such cylinders, each with internal volume of 5 dm^3 and contains 2.58 kg of carbon dioxide. The pressure inside each of the cylinder was found to be 1.0×10^4 kPa at room temperature.

A glass of cola is fizzy because carbon dioxide has been dissolved in it under pressure. When 500 cm^3 of commercial cola drink is poured out from a bottle, the carbon dioxide is gradually released as bubbles of gas. On evolution of 1 dm^3 of

carbon dioxide, the cola went flat, as the concentration of dissolved carbon dioxide decreases to its saturation level of 1.5 g dm^{-3} at room temperature.

(i) Using the ideal gas equation pV = nRT, calculate the pressure the carbon dioxide would exert inside the cylinder at room temperature.

(ii) Suggest why the pressure you calculated in (i) differs from the actual value.

(iii) Calculate the total mass of carbon dioxide dissolved in the 500 cm^3 bottle of cola under pressure.

(iv) Determine the number of 500 cm^3 bottles of cola that could be manufactured using one cylinder of pressurised carbon dioxide.

6(a) Alfred Noble discovered that an explosive could be made from nitroglycerin(a liquid which on its own readily detonates) mixed with a substance called Kieselguhr (diatomaceous earth). The equation for the explosion is:

$$4C_3H_5N_3O_9(l) \longrightarrow 6N_2(g) + 12CO_2(g) + O_2(g) + 10H_2O(g)$$

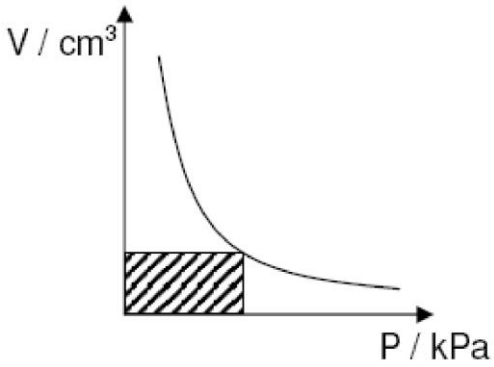
This discovery contributed to his enormous wealth, part of which he used to establish the Nobel Prizes.

(i) Given that the volume of gases produced at 500^0C and 1 atm was 83.0 dm^3. Calculate the mass of nitroglycerin that had decomposed.

(ii) Apart from the great amount of heat energy released, why would this reaction cause extensive damage if it is carried out in a closed area?

(b) Among the gaseous products formed at 500 ^{o}C, Alfred Noble studied the behaviour of these gases. State and explain which of these gaseous products behave most ideally.

(c) The graph below was mapped out by Alfred Noble. He related the behaviour of the most ideal gas at constant temperature.



Given that the area of the shaded region = 5.82 x 104 cm^3 kPa, calculate the pressure exerted by the gas occupying a volume of 112 cm^3.

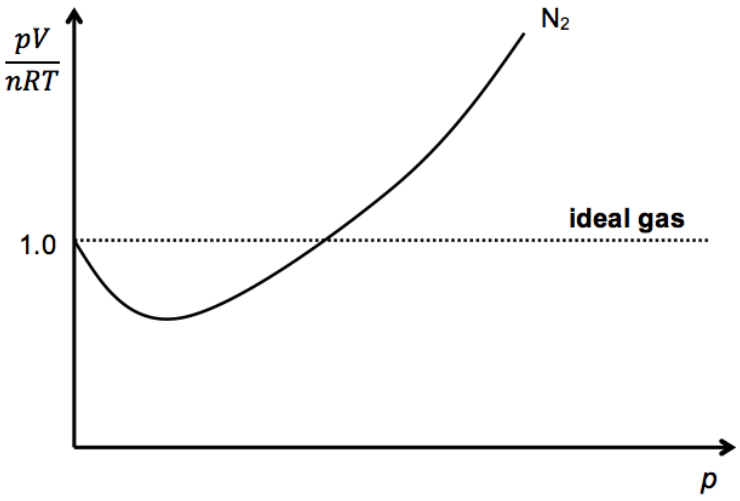
7 Hydrazine, N_2H_4, is a colourless flammable liquid with an ammonia-like odour.

It boils at 114 °C, and gaseous hydrazine can decompose to form hydrogen and nitrogen gases.

(a) Draw the dot-and-cross diagram of N_2H_4.

(b) State the hybridisation of the N atom and suggest what the H–N–H bond angle is in N_2H_4.

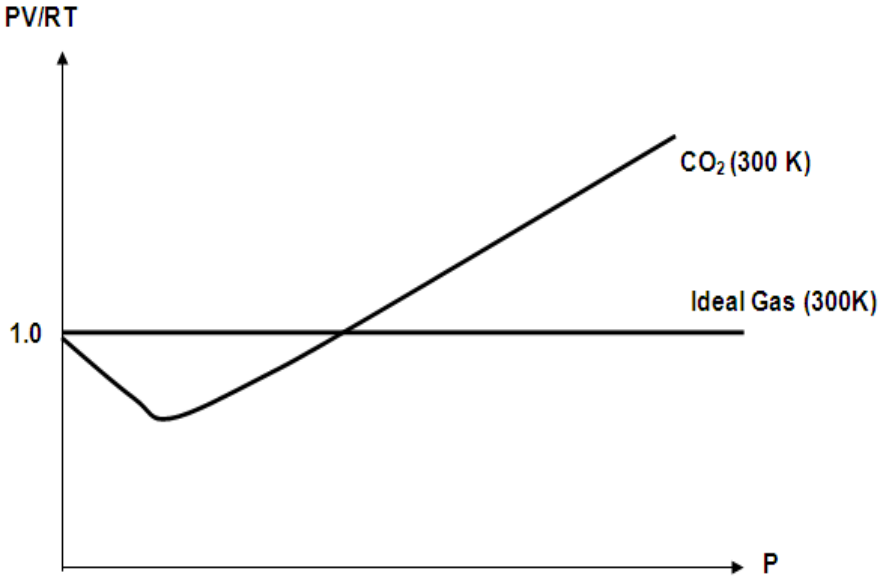
A graph of $\frac{pV}{nRT}$ against p for N_2(g) is given below.



(c) On the axes above, sketch the corresponding graphs for H_2(g) and N_2H_4(g).

8 The plots of pV/RT against p for one mole of an ideal gas and one mole of carbon dioxide gas from the decomposition of magnesium carbonate at 300K are given below.

(i) Show, on the same axes, how one mole of sulfur dioxide, SO_2 will behave at the same temperature of 300K. Label your graph clearly.



(ii) Explain the difference in behavior between carbon dioxide and sulfur dioxide at 300K.

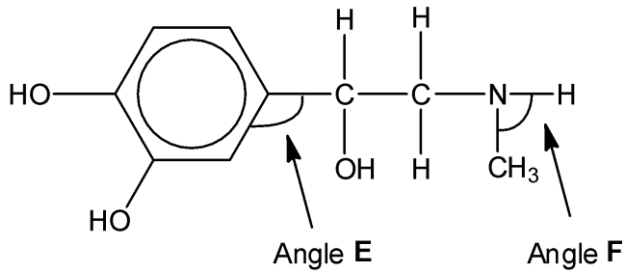
(iii) Explain what happens when sulfur dioxide gas is cooled to 100K. Illustrate your answer clearly on the same axes in **b(i).**

Answers: 5) Ideal Gases

1	**A**	**6**	**C**	**11**	**B**	**16**	**D**	**21**	**B**	**26**	**A**	**31**	**A**
2	**A**	**7**	**A**	**12**	**D**	**17**	**B**	**22**	**C**			**32**	**B**
3	**D**	**8**	**A**	**13**	**C**	**18**	**D**	**23**	**D**	**28**	**B**	**33**	**C**
4	**C**	**9**	**B**	**14**	**C**	**19**	**B**	**24**	**D**	**29**	**B**	**34**	**B**
5	**B**	**10**	**A**	**15**	**B**	**20**	**C**	**25**	**D**	**30**	**B**		

6) Chemical Bonding

1. ***Adrenaline*** is a hormone which, when secreted directly into the blood stream, acts as a stimulant. It has the following structure:



What are the values of angle **E** and angle **F** in a molecule of ***adrenaline***?

	angle ***E***	angle ***F***
A	120°	107°
B	120°	90°
C	109°	107°
D	109°	90°

2. Which of the following has the largest bond angle?

A PH_3 **B** $SiCl_4$ **C** IF_2^- **D** IF_4^-

3. Which of the following pairs contains two species with different shapes?

A NCl_3 and $BrCl_3$ **B** CO_3^{2-} and BBr_3

C $BeCl_2$ and $BrCl_2^-$ **D** SiF_4 and PF_4^+

4. For which of the following pairs does the first species have a smaller bond angle?

1 XeF_4, SeF_2

2 CrO_4^{2-}, SO_3^{2-}

3 BH_3, AsH_3

A 1, 2 and 3 only

B 1 and 2 only

C 2 and 3 only

D 1 only

5. In which set of species is the bond angle of the second species larger than that of the first one?

A NCl_3, H_2O

B BF_4^-, SCl_2

C PCl_6^-, ClO_2^-

D Br_2O, SF_6

---

6. The compound sodium borohydride, $NaBH_4$, is commonly used as a source of the hydride anion for synthetic purposes. Which of the following types of bonding are found in the compound?

1 ionic

2 covalent

3 hydrogen bonding

A 1, 2 and 3 only

B 1 and 2 only

C 2 and 3 only

D 1 only

7. Which of the following pairs of substances have different types of bonding and structure?

A	AlF_3	$AlCl_3$	**B**	Diamond	SiC
C	MgF_2	BaI_2	**D**	PH_3	$SiCl_3H$

8. Which of the following pairs of substances **does not** include a giant structure and a simple molecular structure?

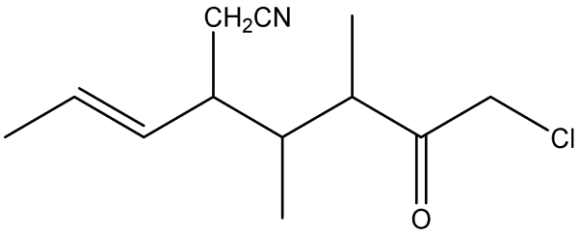
A aluminum and silicon(IV) oxide

B aluminium oxide and aluminium chloride

C silicon and chlorine

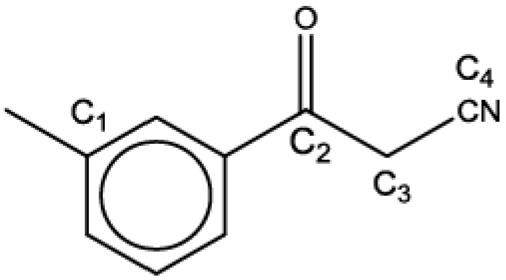
D silicon and silicon(IV) chloride

9. Which of the following options about the structure (as shown below) is correct?



	No of *sp* hybridised C	No of sp^2 hybridised C	No of sp^3 hybridised C
A	1	3	6
B	1	3	8
C	0	4	6
D	0	4	8

10 What is the hybridisation of the various carbon atoms in the following mole



	C_1	C_2	C_3	C_4
A	sp^2	sp^3	sp	sp^2
B	sp^2	sp^2	sp^3	sp
C	sp^2	sp^2	sp	sp^3
D	sp^3	sp^2	sp^3	sp

11

Which of the molecules are polar?

1 $C_6H_5CH_3$ **2** CH_3OCH_3 **3** SF_4

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

12 Which compound is planar?

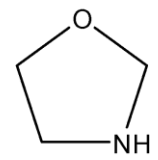
A $C_6H_5CH_3$ **B** $[CuCl_4]^{2-}$ **C** C_6H_{14} **D** O_3

13 Which of the following pairs consists of a planar molecule and a polar molecule?

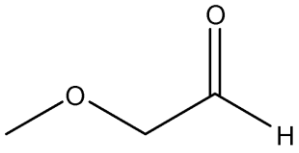
A $C_6H_5CH{=}CH_2$ and PCl_3 **B** $C_6H_5CH{=}CH_2$ and CCl_4

C CH_3COCH_3 and PCl_3 **D** CH_3COCH_3 and CCl_4

14. Which of the following statements best accounts for the difference in the boiling points of oxazolidine and compound **W**?



Oxazolidine



Compound **W**

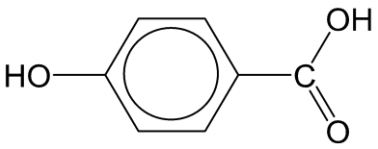
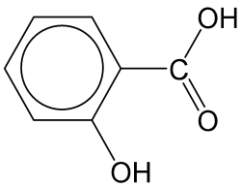
A Oxazolidine is spherical, while **W** is linear.

B Oxygen is more electronegative than nitrogen.

C Oxazolidine has 3 lone pairs of electrons while **W** has 4 lone pairs of electrons.

D Hydrogen bonding is stronger than dipole-dipole interactions.

15 In which of the following pairs does the first substance have a higher melting point than the second?

A	$CH_3CH_2OCH_3$	$CH_3CH_2NHCH_3$
B	$(CH_3)_3CH$	$CH_3CH_2CH_2CH_3$
C	RbC*l*	KC*l*
D		

16. Magnesium oxide and sodium fluoride are isoelectronic.

Which of the following are reasons why the value of the lattice energy of magnesium oxide is more than four times that of sodium fluoride?

1 the higher enthalpy change of hydration of the doubly charged ions
2 the higher electrostatic attraction between the doubly charged ions
3 the shorter interionic distance between the doubly charged ions

A 1, 2 and 3
B 1 and 2 only
C 2 and 3 only
D 1 only

17. Which of the following statements **cannot** be explained by hydrogen bonding?

A Ice has a density lower than water at 0 °C.
B The boiling point of 2-nitrophenol is lower than that of 4-nitrophenol.
C The boiling point of ethanal is higher than dimethyl ether, CH_3OCH_3.
D The relative molecular mass of ethanoic acid is 120 when dissolved in benzene.

18

Which of the following series show an increasing trend in boiling point?

1 Na, Mg, Al

2 PH_3, AsH_3, NH_3

3 BF_3, BCl_3, AlF_3

.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

19. The radius and charge of six ions are given below.

Ion	J^+	L^+	M^{2+}	X^-	Y^-	Z^{2-}
Radius / nm	0.14	0.18	0.15	0.14	0.18	0.15

The ionic solids **MZ**, **JX** and **LY** are of the same lattice type.

Which of the following statements are correct?

1 The melting point of **JX** is higher than that of **LY**.

2 The numerical value of the lattice energy of **MZ** is greater than that of **JX**.

3 The numerical value of the hydration energy of J^+ is smaller than that of M^{2+}.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

Structured Questions

1 [MI/2010] Three compounds, **X**, **Y**, **Z**, have the formula C_3H_8O with boiling points of 10.8 °C, 82.4 °C and **97.4** °C respectively.

(a) Write structural formulas for **X**, **Y** and **Z**.

[3]

(b) Account for your answer in **(a)** based on the molecular structures and the types of intermolecular forces in each.

[3]

(c) A fourth compound with the formula $C_2H_4O_2$, has the same molar mass as the three compounds above and boils at 117.9 ^oC. Propose a structure for this compound and account for its higher boiling point.

[2]

[Total: 8]

2 [AJC 08] The melting points of glycine and two other compounds are given below.

Compound	M_r	Melting Point/^oC
glycine	75	262
1-aminobutane	73	- 49
propanoic acid	74	- 21

Explain the difference in melting point between

I glycine and 1-aminobutane and,

II 1-aminobutane and propanoic acid.

[5]

3(a) [TPJC/07] Boron trifluoride, BF_3, and aluminium fluoride, A/F_3, differ markedly in their physical properties.

Compound	Melting point / ^oC
BF_3	-144
A/F_3	1291

Explain, in terms of structure and bonding, the difference in the melting points of the two compounds. [4]

(b) Ethanoic acid, CH_3CO_2H, is an organic acid. The table below shows the M_r values of the acid obtained in two different solvents.

Solvent	M_r of ethanoic acid
Water	60

Non-aqueous solvent	120

(i) Suggest an explanation for the above observations.

(ii) Draw a diagram to show the type of bonding present when the acid is dissolved in a non-aqueous solvent.

[4]

4 [SRJC/07] An interesting parallel between boron-nitrogen chemistry and carbon chemistry is exemplified by boron nitride, $(BN)_x$. Boron nitride can exist in a hexagonal form, similar to graphite. Hexagonal boron nitride and graphite are both slippery and insoluble in both water and non-polar solvents. However, they have different chemical properties because of the difference in electronegativity between boron and nitrogen. In graphite, the €-electrons are involved in an extended delocalised system and this gives rise to high electrical conductivity. Unlike graphite, hexagonal boron nitride is a poor conductor of electricity.

Layers of hexagonal rings, which share common edges, made up the structure of boron nitride. Within a layer, the **B**-N distances are 0.145 nm. In hexagonal boron nitride, the layers lie over each other and are arranged in such a way that boron and nitrogen atoms alternate in arrangement. This contrasts with the staggered pattern found in graphite. In hexagonal boron nitride, the inter-layer separation is 0.330 nm.

It is suggested that boron nitride can be made an electrical conductor by inserting an alkali metal into its structure. A scientist also discovered that the taste of solution of alkali metal halides depends on the sum of the ionic radii of the ions: i.e. sodium iodide and potassium chloride are salty; rubidium chloride is both salty and bitter; caesium chloride and rubidium bromide are bitter. The taste of the respective alkali metal halides can be inferred from the data below:

Ion	**Radius / nm**	**Ion**	**Radius / nm**
Na^+	0.098	Cl^-	0.181
K^+	0.133	Br^-	0.196

Rb^+	0.148	I^-	0.219
Cs^+	0.167		

Table 1

(a) Describe the structure and bonding in boron nitride.

[2]

(b) Draw the three-dimensional structure of boron nitride, indicating clearly the **B**-N distance *within a layer* and the distance *between two layers*. Label on your structure, *all* the types of interactions present in boron nitride.

[4]

(c) Suggest why hexagonal boron nitride is a poor conductor of electricity.

[1]

(d) Suggest one probable industrial use of boron nitride.

[1]

(e) Use the data in **Table 1** to predict the tastes of sodium bromide and potassium iodide, giving your reasoning.

[4]

5 **A** student was given three solid samples, **X**, **Y** and **Z,** to analyse. The results of the investigations are as follows:

Te	**Observations**
1	When the solids are hit with a hammer, **X** smashes into powder, while **Y** flattens out, and **Z** shatters into many pieces.
2	In their solid states, **Y** is a good conductor of electricity while **X** and **Z** are poor electrical conductors.
3	When the solids are heated at 150 ^oC, **X** sublimes easily, but **Y** and **Z** did not melt.
4	When added to non-polar solvents, **X** dissolves and there is no effect on **Y** and **Z**. However, when water is used, only **Z** dissolves.

(i) Suggest the type of structure in **X**, **Y** and **Z**.

(ii) With reference to bonding in **Y** and **Z**, explain the difference in their electrical conductivity in the solid state.

(iii) With reference to bonding, explain the probable behaviour of **X** and **Z** when dissolved in the respective solvents.

- Solubility of **X** in non-polar solvent:
- Solubility of **Z** in water:

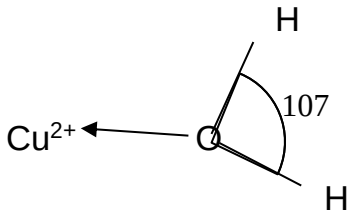
[6]

6 The NO_2 molecule has an unpaired electron on its nitrogen atom but the N_2O_4 molecule does not. Draw a dot–and–cross diagram for NO_2 and suggest a structural formula for N_2O_4, showing its shape and bond angles clearly.

7 [JJC 08] Chalcanthite, known as 'copper flower', is a water soluble copper sulfate mineral. When dissolved, aqueous copper(II) sulfate contains $[Cu(H_2O)_6]^{2+}$ ions.

Water is simple molecular with H•O•H bond angle of 105° in an isolated H_2O molecule.

The diagram below shows part of the $[Cu(H_2O)_6]^{2+}$ ion and the H•O•H bond angle in the water ligand.



(i) Explain why the H–O–H bond angle in the water ligand is 107° while that in an isolated water molecule is 105°.

(ii) State the hybridisation state of the oxygen atom in the water ligand.

[5]

8(a) The table below gives the melting points of some elements of Period 3.

Element	phosphorus	sulfur	chlorine
Melting point /K	317	392	172

In terms of structure and bonding, explain why sulfur has the highest melting point among the three elements.

[2]

(b) A chloride of sulfur was found to contain 47.4% by mass of sulfur and has a boiling point of 136 ^{o}C.

(i) Calculate the empirical formula of this chloride.

(ii) At a pressure of 100 kPa and a temperature of 500 K, 0.100 g of the chloride was found to have a volume of 30.8 cm^3.

Calculate the relative molecular mass of the chloride and hence deduce the molecular formula of the chloride.

(iii) Draw a dot-and-cross diagram for this chloride and state its shape.

[5]

9 At 75 ^{o}C, bromine, Br_2, reacts with aqueous alkali to form Br^- and BrO_3^- ions

(i) Write a balanced equation for the reaction and state the type of reaction undergone by bromine.

(ii) Draw the dot-and-cross diagram of BrO_3^- ion.

Hence, using VSERP theory, predict the shape of BrO_3^- ion.

[4]

10 [VJC 09] This question is about Period 3 elements and their compounds.

(a) A Period 3 oxide, sulfur trioxide dissolves in water to form sulfuric acid, H_2SO_4 which can be converted into peroxodisulfuric acid, $H_2S_2O_8$ via the two-step process below, with chlorosulfonic acid, HSO_3Cl, as an intermediate:

$$H_2SO_4 + HCl \rightarrow HSO_3Cl + H_2O$$

$$H_2O_2 + 2HSO_3Cl \rightarrow H_2S_2O_8 + 2HCl$$

(i) The boiling point of H_2SO_4 is 290 ^{o}C whereas that of HSO_3Cl is 152 ^{o}C. By comparing the structures of the two compounds, explain the difference in the boiling points.

(ii) Give the displayed formula of the $H_2S_2O_8$ molecule, showing the spatial arrangement and estimated bond angle around any one sulfur atom and one oxygen atom.

[5]

(b)(i) Define the term *lattice energy.*

(ii) State and explain whether the oxide or fluoride ion is smaller.

(iii) Refractories are heat-resistant compounds used to line furnaces. Which compound, magnesium oxide or magnesium fluoride, is more suitable for use as a refractory? Explain your answer.

[5]

(c) Magnesium fluoride, differs in its bonding from that of sulfur difluoride, SF_2, as reflected by the difference in their physical properties.

(i) Explain why the difference in bonding arises.

(ii) Suggest one difference in physical properties that is likely to ensue from the difference in bonding.

(iii) The compound SF_2 readily reacts with fluorine to give SF_4. Suggest why MgF_2 does not react with more fluorine to give MgF_4.

[4]

11(a) [TPJC 08] Describe the shapes of, and state the bond angles in, the molecules of hydrogen sulfide, H_2S, and sulfur dioxide, SO_2. Briefly explain your answers in terms of the number and types of electron pairs around the central atom.

[4]

(b) The sulfur atom in SO_2 is joined to the oxygen atom by a ***α-bond*** and a ***π-bond***. Explain, with the aid of diagrams, what is meant by the terms in *italics*.

[3]

(c) Carbon dioxide sublimes at –78 ^{o}C whereas the boiling point of sulfur dioxide is –10 ^{o}C. Explain this observation in terms of bonding and structure.

[3]

12 [AJC/07] Magnesium chloride, $MgCl_2$, is a solid with a high melting point (714 °**C**) whereas aluminium chloride, $AlCl_3$, sublimes at 178 °C.

(i) Relate the above data to the structure and bonding in each of these chlorides.

(ii) Draw dot-and-cross diagrams to show the different electron arrangements in the two compounds.

(iii) $AlCl_3$ can react with chlorine gas as shown in the equation below.

$$AlCl_3 + Cl_2 \longrightarrow AlCl_4^- + Cl^+$$

Draw a diagram to illustrate the shape of and bonding in $AlCl_4^-$ and explain why aluminium chloride can react with chlorine gas.

[7]

13 Hydrogen sulfide, H_2S, smells of rotten eggs and is highly poisonous.

1 mole of hydrogen sulfide can react with 2 moles of boron trifluoride, BF_3, to form compound **C** only.

(i) Draw a diagram to illustrate the shape of and bonding in **C**.

(ii) Gaseous hydrogen sulfide is appreciably soluble in water to give a resulting solution that turns universal indicator yellow. Write an equation, with state symbols, to illustrate the solubility of hydrogen sulfide with water.

[3]

14 [RIJC 2010] The structure of methane, CH_4, and ethyne, C_2H_2, may be described in terms of the hybridisation of orbitals.

(a) State the type(s) of hybridisation present in the methane and ethyne molecules.

[2]

(b) Sketch the shapes of the hybrid orbitals around one carbon atom in the ethyne molecule.

[1]

(c) Methane reacts with $HF–SbF_5$, a superacid to form an intermediate, $[CH_5]^+[SbF_6]^-$. The intermediate immediately decomposes to form $[CH_3]^+[SbF_6]^-$ and hydrogen gas.

$$CH_4 + HF\text{-}SbF_5 \rightarrow [CH_5]^+ [SbF_6]^- \rightarrow [CH_3]^+ [SbF_6]^- + H_2$$

(d) Explain why the intermediate formed is not stable.

[1]

(e) State the shapes and bond angles of $[CH_3]^+$ and $[SbF_6]^-$.

[4]

2,2-dimethylpropane reacts similarly with $HF–SbF_5$ to form an ionic salt with negative ion $[SbF_6]^-$ and a gas **A** which contains 75 % by mass of **C**.

(f) Deduce the structural formulae of gas **A** and the positive ion of the salt.

[4]

[Total: 12]

15 When solid aluminium chloride is heated above 180 ^oC, a vapour is formed which has M_r = 267. When this vapour is heated above 800 ^oC, the vapour has M_r = 133.5.

(a)(i) Draw the displayed formula of the vapour at 180 ^oC, to show its bonding. State the shape of the molecule.

(ii) The average M_r of the vapour is found to be 240 at 600 ^oC. Determine the molar composition of the vapour.

[2]

(b) Chlorine forms an oxide, Cl_2O_6, which exists as singly charged ions in the solid state. The oxidation states of chlorine are +5 and +7 in the cation and anion respectively. Suggest the formulae of the ions and draw the dot-and-cross diagram for the **cation**.

[2]

16 [VJC 09] Potassium oxide, K_2O and potassium peroxide, K_2O_2 are two compounds that can be produced when potassium is heated in oxygen under suitable conditions.

(a) Both K_2O and K_2O_2 are ionic compounds containing one anion each.

(i) Write the formulae of the anions in K_2O and K_2O_2.

(ii) Predict the relative magnitudes of the melting points of K_2O and K_2O_2. Explain your answer.

[3]

(b) When water is added to potassium peroxide, K_2O_2, potassium hydroxide and hydrogen peroxide are produced.

(i) Write an equation for the reaction that has occurred.

(ii) The following data refer to the boiling points of hydrogen peroxide, hydrazine and hydrogen sulfide.

compound	formula	boiling point / °C
hydrogen peroxide	H_2O_2	150
hydrazine	N_2H_4	114
hydrogen sulfide	H_2S	–61

Relate these boiling points to the bonding in each of these compounds.

(iii) Draw and name the shape of hydrogen sulfide, H_2S.

(iv) Draw a diagram to show how hydrazine is miscible with water.

[7]

17 [DHS09] This question is about compounds commonly used as chlorinating agents in organic synthesis such as PCl_3, PCl_5 and $SOCl_2$.

(a) Draw the Lewis structures of PCl_5 and $SOCl_2$, state their shapes.

[4]

(b) Predict the solubility of PCl_5 in CCl_4. Support your prediction with appropriate explanation.

[2]

18 [SAJC 09]

(a)(i) Carbon and silicon are Group 14 elements.

Explain why the first ionisation energy of carbon is higher than silicon.

(i) Carbon vapourises at about 4027 °C at 101 kPa. Calculate the volume of gaseous carbon formed when 24 g of carbon is vapourised.

(iii) Sketch a labelled graph to show how the volume of 24 g of gaseous carbon changes when heated from 4027 °C to 5000 °C.

(iv) Explain why gaseous carbon behaves more ideally as compared to nitrogen gas.

[6]

(b) Cyanamide, NH_2CN, is widely used in synthetic organic chemistry and is produced by the hydrolysis of calcium cyanamide, CaNCN.

(i) Draw dot-and-cross diagrams to illustrate the bonding in these two compounds. Hence, predict the shape of cyanamide, NH_2CN.

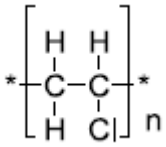
(ii) Describe the structure and bonding present in each of the cyanamide compounds, and hence comment on their melting points.

(ii) Suggest, with brief reasoning, whether each of the cyanamide compounds is expected to be soluble in water.

[8]

[Total: 14]

19 Electric cable is made of copper wire surrounded by poly(chloroethene) which is also called polyvinyl chloride, PVC. PVC has the following formula:



Another type of cable used in fire alarm systems has the copper wire surrounded by solid magnesium oxide, which acts as an insulator. The whole cable is encased in a flexible copper tube.

(a) Describe the bonding in copper, and hence explain how it can be drawn into wires.

(b) Describe the two types of bonding present in PVC.

(c) Suggest one reason why magnesium oxide is preferred to PVC as an insulator in fire alarm cabling.

(d) Suggest why magnesium oxide needs to be encased in copper tube.

20 (NYJC 12) The standard enthalpy change of fusion (ΔH^{θ}_{fus}) is the energy required to convert one mole of a substance in the solid state to the liquid state under standard pressure. The table below shows numerical values of standard enthalpy change of fusion for the respective elements:

Element	ΔH^{θ}_{fus} / kJ mol^{-1}
Na	2.60
Al	10.7
Si	50.2
C*l*	6.41

Explain, in terms of structure and bonding, the difference in the ΔH^{θ}_{fus} between:

(i) Si and C*l*

(ii) Na and A*l*

21 (DHS 13) The data for some liquids are given below:

	M_r	**boiling point/ °C**	**density / g cm^{-3}**
ethanol	46	79	0.79
cyclohexane	84	81	0.78
hexane	86	69	0.66

(a) **(i)** Suggest why the boiling point of cyclohexane differs from that of hexane.

(ii) Suggest why the density of ethanol differs from that of hexane.

(b) Equal volumes of 1.0 mol samples of the following liquids were mixed and the following results were obtained in the following table.

Mixture of liquids	**Heat change**
Hexane mixed with cyclohexane	No heat absorbed or given out
Hexane mixed with ethanol	?

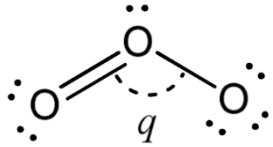
(i) Explain why there was no heat change on mixing hexane and cyclohexane.

(ii) Predict, with reasoning, the heat change on mixing hexane and ethanol.

22 (RV 13) The melting points of NiS and H_2S are as shown below. Explain the difference in their boiling points in term of structure and bonding.

Compound	Melting point / °C
NiS	797
H_2S	– 82

23 (VJC 13) Ozone, O_3, has the following structure.



ozone

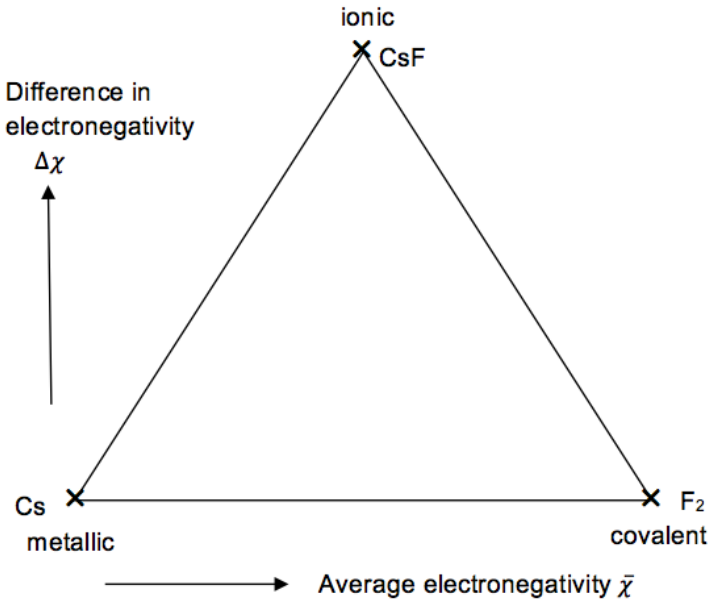
Bond type	Bond length / Å
O–O	1.48
O=O	1.21

(i) The separation between any two adjacent oxygen atoms in ozone is 1.28 Å.
Explain this bond length with suitable chemical structure(s).

(ii) Suggest a value for angle *q*.

24 (HCI 13) The type of bonding present in a binary compound can be predicted from the electronegativities of the elements involved. This can be shown on a van Arkel-Ketelaar triangle, which plots the difference in electronegativity, $\Delta\chi$, on the y-axis against the average electronegativity of the two elements, $\underline{\chi}$, on the x-axis.

In this triangle, the three corners represent the extremes of metallic, ionic, and covalent bonding, with caesium (Cs), caesium fluoride (CsF), and fluorine (F_2) in these corners.



(i) State the type of bonding in sodium chloride (NaC*l*), aluminium chloride (A*l*C*l*$_3$), and titanium-aluminium alloy (Ti-A*l*).

(ii) Mark the approximate positions of NaC*l*, A*l*C*l*$_3$, and Ti-A*l* with crosses (x) in the van Arkel-Ketelaar triangle above, and label the crosses accordingly.

25 (NJC 14) The simplest chemical reactions are those that occur in the gas phase in a single step, such as the transfer of a chlorine atom from C*l*NO$_2$ to NO.

$$ClNO_2(g) + NO(g) \rightleftharpoons NO_2(g) + ClNO(g)$$

(i) Draw the dot-and-cross diagrams of NO and C*l*NO. Suggest the shape and bond angle of C*l*NO.

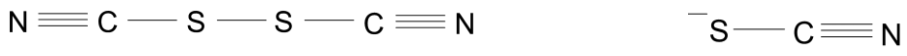
(ii) Hence, suggest why the formation of C*l*NO from NO is favoured.

26 (RV 14) Explain, in terms of bonding and structure, the following observations.

(i) Copper has a higher melting point than calcium.

(ii) Copper conducts electricity in solid state while CuO does not.

27 (VJC 14) Pseudohalogens are a class of compounds that show chemical reactivity similar to halogens. An example of a pseudohalogen and its anion is $(SCN)_2$ and SCN^- respectively.



thiocyanogen thiocyanate

(i) Draw the shape of the hybrid orbitals around one of the carbon atoms in thiocyanogen.

(ii) Draw a diagram to illustrate the shape of thiocyanogen, indicating clearly the covalent bonds and lone pairs of electrons.

(iii) Thiocyanogen can be synthesised by reacting sodium thiocyanate, NaSCN, with chlorine gas. Write an equation for the formation of thiocyanogen.

28 (VJC 14)

(i) Draw the dot-and-cross diagram of ozone, O_3.

(ii) Oxygen–oxygen bond lengths of 3 species are given in the table below.

Species	Oxygen–oxygen bond length / nm
O_3	0.128
O_2^{2-}	0.149
O_2	0.120

With the aid of suitable structures, explain why the two oxygen-oxygen bonds in the ozone molecule have the same bond length which are very different from the peroxide ion and oxygen molecule.

Answers: 6) Chemical Bonding

1	**A**	**5**	**C**	**9**	**B**	**13**	**A**	**17**	**C**
2	**C**	**6**	**B**	**10**	**B**	**14**	**D**	**18**	**A**
3	**A**	**7**	**A**	**11**	**C**	**15**	**D**	**19**	**A**
4	**D**	**8**	**A**	**12**	**D**	**16**	**C**		

7) Chemical Energetics: Thermochemistry and Thermodynamics

MCQ

1 The reduction of the iron ore, Fe_2O_3 by carbon monoxide is an exothermic reaction which occurs in the upper part of the furnace.

$$Fe_2O_3(s) + 3CO(g) \longrightarrow 2Fe(s) + 3CO_2(g) \qquad \Delta H = -27 \text{ kJ mol}^{-1}$$

Given that ΔH_f^θ for $CO_2(g)$ = –394 kJ mol^{-1} and ΔH_f^θ for $CO(g)$ = –110 kJ mol^{-1}, what is the standard enthalpy change of formation of the iron ore?

A +257 kJ mol^{-1} **B** +825 kJ mol^{-1} **C** –257 kJ mol^{-1} **D** –825 kJ mol^{-1}

2 Given the following information:

lattice energy of magnesium chloride = −2526 kJ mol^{-1}

enthalpy change of solution of magnesium chloride = −132 kJ mol^{-1}

enthalpy change of hydration of chloride ion = −384 kJ mol^{-1}

what is the enthalpy change of hydration of the magnesium ion?

A −516 kJ mol^{-1} **B**−1890 kJ mol^{-1} **C** −2274 kJ mol^{-1} **D** +2778 kJ mol^{-1}

3 Given the following enthalpy changes:

$$C_6H_4(OH)_2(aq) \rightarrow C_6H_4O_2(aq) + H_2(g) \qquad \Delta H = +177 \text{ kJ mol}^{-1}$$

$$H_2O_2(aq) \rightarrow H_2O(l) + \tfrac{1}{2} O_2(g) \qquad \Delta H = -95 \text{ kJ mol}^{-1}$$

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

what is the enthalpy change for the following reaction?

$$C_6H_4(OH)_2(aq) + H_2O_2(aq) \rightarrow C_6H_4O_2(aq) + 2H_2O(l)$$

A −490 kJ mol^{-1} **B** −368 kJ mol^{-1} **C** −204 kJ mol^{-1} **D** −82 kJ mol^{-1}

4 Consider the following enthalpy changes for the two reactions of ethyne C_2H_2.

$C_2H_2(g) + H_2(g) \longrightarrow C_2H_4(g)$ ΔH_1

$C_2H_2(g) + 2H_2(g) \longrightarrow C_2H_6(g)$ ΔH_2

Which of the expressions represents ΔH_3?

$C_2H_4(g) + H_2(g) \longrightarrow C_2H_6(g)$ ΔH_3

A $\Delta H_2 - \Delta H_1$ **B** $\Delta H_1 - \Delta H_2$ **C** $-\Delta H_1 - \Delta H_2$ **D** $\Delta H_1 + \Delta H_2$

5 You are given the following standard enthalpy changes:

	ΔH/ kJ mol^{-1}
$C(graphite) + 2H_2(g) \rightarrow CH_4(g)$	−75
$H_2(g) \rightarrow 2H(g)$	+436

Using any relevant data from the *Data Booklet*, what is the standard enthalpy change of atomisation of graphite?

A +693 kJ mol^{-1} **B** +1129 kJ mol^{-1} **C** −2151 kJ mol^{-1} **D** −2587 kJ mol^{-1}

6 The enthalpy change of formation of $H_2S(g)$ and $H_2O(l)$ are −21 kJ mol^{-1} and −286 kJ mol^{-1} respectively and the enthalpy change of vaporisation of water is +41 kJ mol^{-1}.

What is the enthalpy change of reaction for the following process?

$H_2S(g) + ½O_2(g) \rightarrow H_2O(g) + S(s)$

A −224 kJ mol^{-1} **B** −266 kJ mol^{-1} **C** −306 kJ mol^{-1} **D** −348 kJ mol^{-1}

7 The major natural source of fluorine is the mineral fluorspar, which is mainly calcium fluoride, CaF_2. The first stage in liberating F_2 is to grind the compound up and react it with concentrated sulfuric acid. The products are hydrogen fluoride and calcium sulfate.

$$CaF_2 + H_2SO_4 \rightarrow 2HF + CaSO_4$$

What is the enthalpy change of the above reaction (in kJ mol^{-1}), given the following information?

compound	**ΔH_f / kJ mol^{-1}**
CaF_2	−1220
H_2SO_4	−814
HF	−271
$CaSO_4$	−1434

A +58 **B** –58 **C** +329 **D** –329

8 Phosphorus can react with chlorine to form phosphorus pentachloride and phosphorus trichloride. Some thermochemical data related to these compounds are shown below:

Enthalpy change of atomisation of phosphorus	+315 kJ mol^{1}
Enthalpy change of atomisation of chlorine	+122 kJ mol^{1}
Enthalpy change of formation of PCl_5	−440 kJ mol^{1}

What is the bond energy for the P–C*l* bond in PCl_5?

A 273 kJ mol^{-1} **B** 485 kJ mol^{-1} **C** 925 kJ mol^{-1} **D** 1365 kJ mol^{-1}

9 The heat liberated in 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 heat liberated in 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}

10 The table below gives the standard enthalpy change of hydrogenation of three compounds to form cyclohexane:

Compound	$\Delta H_{hydrogenation}$ /kJ mol^{-1}
benzene	−206
1,3,5−cyclohexatriene	−360
cyclohexene	−120

Which of the following statements is correct?

A 1,3,5−cyclohexatriene is more stable than benzene.

B Benzene has a lower energy content than 1,3,5−cyclohexatriene.

C The C=C bond energy in cyclohexene is weaker than that in 1,3,5−cyclohexatriene.

D The enthalpy change of atomisation of benzene is smaller than that of 1,3,5−cyclohexatriene.

11 For which of the following ions is the enthalpy change of hydration likely to be the most exothermic?

	Ion	*ionic radius/nm*
A	W^-	0.181
B	X^+	0.169
C	Y^{2+}	0.065
D	Z^{2+}	0.135

12 The enthalpy change of reaction between calcium and water, ΔH_r, can be measured in the laboratory.

$$Ca(s) + 2H_2O(l) \rightarrow Ca(OH)_2(s) + H_2(g); \quad \Delta H_r$$

What other data is needed in order to calculate the enthalpy change of formation of $Ca(OH)_2(s)$?

A enthalpy change of atomisation of calcium

B enthalpy change of combustion of hydrogen

C first and second ionisation energies of calcium

D lattice energy of calcium hydroxide

13 The lattice energies of rubidium fluoride, RbF, and caesium chloride, CsC*l*, are −760 kJ mol^{-1} and −650 kJ mol^{-1} respectively.

What is the lattice energy of caesium fluoride, CsF, likely to be?

A −520 kJ mol^{-1} **B** −580 kJ mol^{-1} **C** −720 kJ mol^{-1} **D** −920 kJ mol^{-1}

14 Strontium hydride is an ionic compound. The standard enthalpy changes for the following two reactions related to its constituent elements and ions are known.

$$Sr(s) \downarrow\rightarrow Sr^{2+}(g) + 2e^-$$

$$H_2(g) + 2e^- \downarrow\rightarrow 2H^-(g)$$

What additional data is needed to determine the lattice energy of strontium hydride?

A first ionisation energy of hydrogen

B standard enthalpy change of atomisation of hydrogen

C standard enthalpy change of atomisation of strontium

D standard enthalpy change of formation of strontium hydride

15 *The use of the Data Booklet is relevant to this question.*

Aluminium oxide is a stable ionic compound.

$$Al(s) \rightarrow Al^{3+}(g) + 3e \qquad \Delta H = +5467 \text{ kJ mol}^{-1}$$

$$\tfrac{1}{2} O_2(g) + 2e \rightarrow O^{2-}(g) \qquad \Delta H = +897 \text{ kJ mol}^{-1}$$

Which of the following statements is ***false***?

A The enthalpy change of atomisation of aluminium is +330 kJ mol^{-1}.

B The enthalpy change of formation of aluminium oxide is –13625 kJ mol^{-1}.

C The lattice energy of aluminium oxide is more negative than the enthalpy change of formation of aluminium oxide.

D The sum of the first and second electron affinities of oxygen is +649 kJ mol^{-1}.

16 The standard enthalpy changes of formation of HC*l* and HI are –92 kJ mol^{-1} and +26 kJ mol^{-1} respectively.

Which statement is best explains the difference?

A The bond energy of HI is smaller than that of HC*l*.

B The bond energy of I_2 is smaller than that of Cl_2.

C Chlorine is more electronegative than iodine.

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

17 Graphite and diamond are two allotropes of carbon. Although graphite is thermodynamically more stable than diamond at 25 °C and 1 bar, a piece of diamond will not transform into graphite even over a period of one million years.

Which of the following correctly explains this observation?

A ∆G for the reaction *diamond* → *graphite* is greater than zero.

B ⊗H for the reaction *diamond* → *graphite* is greater than zero.

C The reverse reaction *graphite* → *diamond* proceeds quickly.

D The reaction *diamond* → *graphite* has a large activation energy.

18 Which equation correctly defines the enthalpy change of formation of a compound?

A $H_{2(}g) + O_{2(}g) \rightarrow H_2O_2(l)$

B $N_2H_{4(}l) + O_{2(}g) \rightarrow N_{2(}g) + 2H_2O(g)$

C $CO(g) + ½\ O_{2(}g) \rightarrow CO_2(g)$

D $Na(g) + Cl(g) \rightarrow NaCl(g)$

19 The standard enthalpy of formation, $\Delta H^{\ominus}_f$ for $O_3(g)$ is +142 kJ mol^{-1}.

Which statement about ozone is **incorrect**?

A Decreasing the temperature will increase the amount of O_2 gas in an equilibrium mixture containing only $O_2(g)$ and $O_3(g)$.

B Decreasing the pressure will increase the amount of O_2 gas in an equilibrium mixture containing only $O_2(g)$ and $O_3(g)$.

C O_3 is energetically less stable than oxygen.

D The activation energy for converting O_3 to O_2 is higher than that for O_2 to O_3.

20 When 20 cm^3 of 1 mol dm^{-3} NaOH(aq) was added to 10 cm^3 of 1 mol dm^{-3} HCl(aq), a temperature rise of 5 ^{o}C was obtained.

Assuming that the density of each solution = 1 g cm^{-3} and that no heat is lost to the surroundings, which expression gives the enthalpy change of neutralisation? (Specific heat capacity of the mixed solutions = 4.2 J g^{-1}K^{-1})

A $-\dfrac{30 \times 4.2 \times 5 \times 1000}{10}$ J

B $-\dfrac{20 \times 4.2 \times 5 \times 1000}{10}$ J

C $-\dfrac{10 \times 4.2 \times 5 \times 1000}{30}$ J

D $-\dfrac{30 \times 4.2 \times 5 \times 1000}{30}$ J

21 A student dissolved 8.4 g of sodium fluoride in 250 g of water.

Given the following thermodynamic data,

Lattice energy of NaF	−918 kJ mol^{-1}
Enthalpy change of hydration of F−	−457 kJ mol^{-1}
Enthalpy change of hydration of Na^+	−390 kJ mol^{-1}

What would be the initial temperature of the water if the final temperature of the solution is 20.00 °C? [Assume specific heat capacity of NaF(aq) = 4.2 J g^{-1} K^{-1}.]

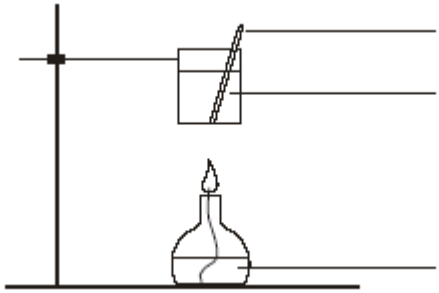
A 6.48 °C **B** 19.99 °C **C** 20.01 °C **D** 33.52 °C

22 When 0.2 mol of zinc dust was added to 250 cm^3 of 1.0 mol dm^{-3} aqueous copper(II) sulfate (in excess), the temperature of the solution rose by 15°C. The specific heat capacity of the final solution is 4.20 J g^{-1} K^{-1}.

What is the enthalpy change for the reaction $Cu^{2+} + Zn \rightarrow Zn^{2+} + Cu$?

A −78.8 kJ mol^{-1} **B** −39.4 kJ mol^{-1} **C** −15.8 kJ mol^{-1} **D** −3.15 kJ mol^{-1}

23 A student used the apparatus below to heat a can containing 300 g of water.



The following data were recorded:

Mass of propan-1-ol burnt = *m* g

Change in temperature of water = Δ*T* °C

You are given:

relative molecular mass of propan-1-ol = 60.0

enthalpy change of combustion of propan-1-ol = −2021 kJ mol^{-1}

specific heat capacity of water = *c* kJ kg^{-1} K^{-1}

What is the efficiency of this heating process?

A $\frac{m \times 2021 \times 1000}{300 \times c \times \Delta T \times 60.0} \times 100\%$ **B** $\frac{m \times c \times \Delta T \times 60.0}{300 \times 2021 \times 1000} \times 100\%$

C $\frac{300 \times c \times \Delta T \times 60.0}{m \times 2021} \times 100\%$ **D** $\frac{300 \times c \times \Delta T \times 60.0}{m \times 2021 \times 1000} \times 100\%$

24 When 25 cm^3 of 1 mol dm^{-3} potassium hydroxide is neutralised with an equal volume of 1.0 mol dm^3 hydrochloric acid, the temperature of the mixture rose by 6.8 °C.

What will be the temperature change if 50 cm^3 of 0.5 mol dm^{-3} potassium hydroxide is neutralised with an equal volume of 0.5 mol dm^{-3} hydrochloric acid? (Assume heat loss to be negligible in each case.)

A 1.7 °C **B** 3.4 °C **C** 6.8 °C **D** 13.6 °C

25 One mole of N_2O gas is completely decomposed in a sealed container, at a constant pressure, according to the following equation:

$$N_2O(g) \longrightarrow N_2(g) + ½O_2(g)$$

Which of the following changes, while keeping the pressure constant, is least likely to increase the entropy of this system?

A increase the temperature **B** remove oxygen from the system

C add an inert gas into the system **D** add more N_2O gas into the system

26 Ozone in the upper atmosphere reduces damaging ultraviolet light from reaching the Earth's surface. The following equation shows the formation of ozone:

$$3O_2(g) \rightarrow 2O_3(g)$$

Given: $\Delta H^{\ominus}_f(O_3)$ = +142.67 kJ mol^{-1} and $\Delta S^{\ominus}(O_3)$ = –68.7 J K^{-1} mol^{-1}, what is the value of $\Delta G^{\ominus}_f(O_3)$ at 298 K and 1 bar?

A +122 kJ mol^{-1} **B** +163 kJ mol^{-1} **C** –20330 kJ mol^{-1} **D** +20615 kJ mol^{-1}

27 For which process is the enthalpy change always negative?

A burning an element in oxygen **B** dissolving a compound in water

C forming an ion from an atom **D** forming a compound from its elements

28 Which equation below represents a reaction that is spontaneous at all temperatures?

A	**P**(s) → **Q**(s) + **R**(g)	ΔH >0, ΔS : + spontaneous at high T
B	2**L**(g) + **M**(g) → 2**N**(g)	ΔH <0, ΔS : − spontaneous at low T
C	**S**(g) → 2**T**(g)	ΔH <0, ΔS : + spontaneous at all T
D	**A**(g) + **B**(g) → **C**(g)	ΔH >0, ΔS : − *non*-spontaneous at all T

29 *The use of the Data Booklet is relevant to this question.*

Dinitrogen oxide, N=N=O, burns in ethyne, C_2H_2, in the gaseous phase to produce water vapour, carbon dioxide and nitrogen gases as the only products.

$5\ N_2O(g) + C_2H_2(g) \rightarrow H_2O(g) + 2\ CO_2(g) + 5\ N_2(g)$ $\quad \Delta H = -1668\ kJ\ mol^{-1}$

Assuming N=N bond energy in dinitrogen oxide is +418 kJ mol^{-1}, what is the nitrogen-oxygen bond energy in dinitrogen oxide in kJ mol^{-1}?

A 382 **B** 594 **C** 686 **D** 1350

30 *The use of the Data Booklet is relevant to this question.*

Propanone is a common industrial solvent, which can be converted to propan-2-ol by hydrogenation.

$$CH_3COCH_3(g) + H_2(g) \rightarrow CH_3CH(OH)CH_3(g)$$

What is the enthalpy change for the hydrogenation reaction?

A + 54 kJ mol^{-1} **B** –54 kJ mol^{-1} **C** + 434 kJ mol^{-1} **D** –434 kJ mol^{-1}

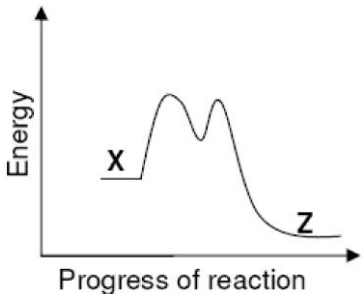
31 The conversion of compound **X** into **Z** was exothermic and proceeded by two-step mechanism where **Y** was the intermediate. The steps involved were:

Step 1: **X** ⟶ **Y** slow

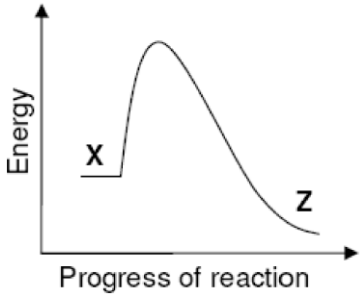
Step 2: **Y** ⟶ **Z** fast

Which is the energy level diagram for the reaction?

A



B



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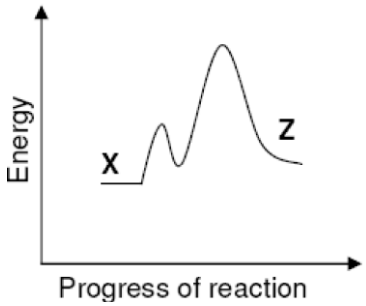

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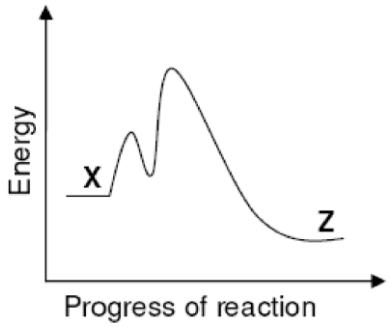
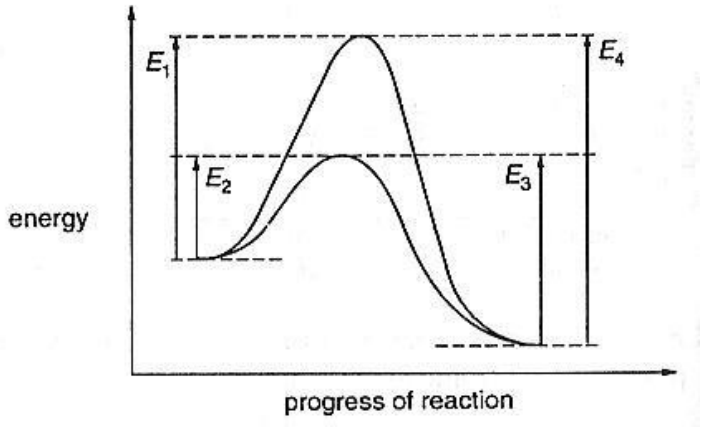


32 The energy diagram represents the reaction occurring with and without a catalyst.



Which of the following statements is correct?

A E_4 is the activation energy for the reverse catalysed reaction.

B The forward reaction, with catalyst, is endothermic.

C The enthalpy change of reaction is ($E_2 - E_3$).

D The enthalpy change of reaction is reduced by using a catalyst.

34 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 values of the following data are needed in order to perform the calculation?

A $\Delta H^{\ominus}_{c}$ (C), $\Delta H^{\ominus}_{c}$ (H_2), $\Delta H^{\ominus}_{c}$ (CH_4)

B $\Delta H^{\ominus}_{at}$ (C), $\Delta H^{\ominus}_{at}$ (H_2), $\Delta H^{\ominus}_{f}$ (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.

35 The enthalpy changes of formation of gaseous ethene and ethane are +52 kJ mol^{-1} and –85 kJ mol^{-1} respectively at 298 K.

Consider the following reaction:

$$C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$$

Which statement is true?

A $\Delta S_{reaction}$ is positive.

B The enthalpy change of the reaction is –137 kJ mol^{-1}.

C If ethene is converted to ethane via four stages instead of one as shown above, the enthalpy change of the reaction will be less exothermic.

D The addition of a catalyst will cause the enthalpy change of reaction to be more exothermic.

36 When water vaporises, 51.6 kJ mol^{-1} of heat enthalpy is required. What is the entropy change when 54 g of water boils at 1 bar?

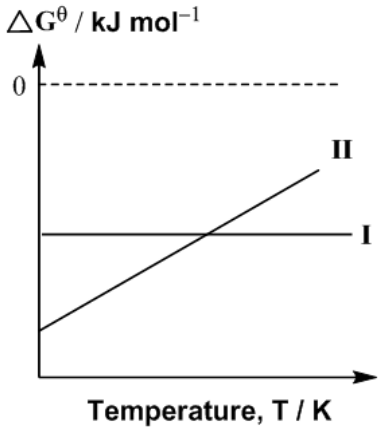
A + 138 J K^{-1}

B + 415 J K^{-1}

C + 516 J K^{-1}

D + 1548 J K^{-1}

37 Ellingham diagrams are plots of ΔG against temperature, T. An Ellingham diagram for two reactions involving the oxidation of S and SO_2 is as shown below:



Given that, ΔG = ΔH - TΔS, which of the following correctly shows the two reactions corresponding to reactions **I** and **II** in the above Ellingham diagram?

	I	**II**
A	$2S + 3O_2 \rightarrow 2SO_3$	$2SO_2 + O_2 \rightarrow 2SO_3$
B	$2SO_2 + O_2 \rightarrow 2SO_3$	$S + O_2 \rightarrow SO_2$
C	$S + O_2 \rightarrow SO_2$	$2S + 3O_2 \rightarrow 2SO_3$
D	$2SO_2 + O_2 \rightarrow 2SO_3$	$2S + 3O_2 \rightarrow 2SO_3$

38 In a blast furnace, carbon or carbon monoxide can be used to reduce iron(III) oxide.

$Fe_2O_3(s) + 3/2\ C(s) \rightarrow 2Fe(s) + 3/2\ CO_2(g)$ $\Delta H = +234$ kJ mol^{-1} (1)

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$ $\Delta H = -24.8$ kJ mol^{-1} (2)

Carbon monoxide can be formed by the following reaction.

$C(s) + CO_2(g) \rightarrow 2CO(g)$ ΔH_1

What is the value of ΔH_1?

A +86.3 kJ mol^{-1}

B +139.5 kJ mol^{-1}

C +172.5 kJ mol^{-1}

D +258.8 kJ mol^{-1}

39 Which of the following reactions shows a positive change in entropy?
[Assume that all measurements are taken at 298 K and 1 bar pressure.]

A $N_2O(g) \rightarrow N_2(g) + ½O_2(g)$

B $Fe(s) + S(s) \rightarrow FeS(s)$

C $Ca(s) + ½O_2(g) \rightarrow CaO(s)$

D $C_3H_6(g) + H_2(g) \rightarrow C_3H_8(g)$

40 1600 J of energy is required to vaporise 40 cm^3 of a pure liquid **Q**. The ΔS of vaporisation of liquid **Q** is found to be +100 J mol^{-1} K^{-1}. Given that the molar density of **Q** is 1.0 mol dm^{-3}, what is the boiling point of liquid **Q** at 1 bar?

A 200 K

B 400 K

C 600 K

D Cannot be determined from the above data

41 For a chemical reaction,

$$A \rightleftharpoons B \qquad \Delta H = +67.0 \text{ kJ mol}^{-1}$$

$$\Delta S = -131 \text{ J K}^{-1} \text{ mol}^{-1}$$

Which of the following statement is correct?

A The reaction will be spontaneous when temperature is less than –511 K.

B The reaction will be spontaneous when temperature is greater than –511 K.

C The reaction will be spontaneous at any temperature.

D The reaction will not be spontaneous at any temperature.

42 Which one of the following equations represents a reaction that is feasible at all temperatures?

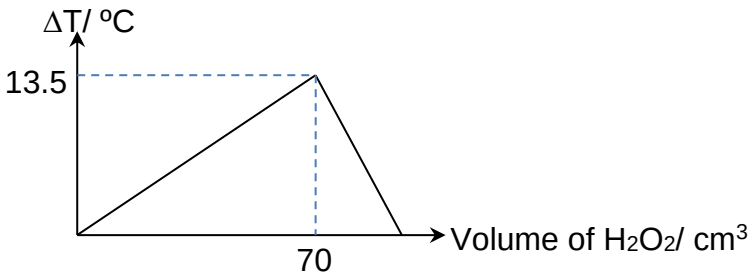
A **W**(s) → **X**(s) + **Y**(g) $\Delta H > 0$

B 2**T**(g) + 3**U**(g) → 4**V**(g) $\Delta H < 0$

C **R**(g) → 2**Q**(g) $\Delta H < 0$

D **K**(g) + **L**(g) → **M**(g) $\Delta H > 0$

43 Varying volumes of H_2O_2 of an unknown concentration was reacted with varying volumes of 0.50 mol dm^{-3} acidified potassium manganate(VII) in a series of experiments, to produce the graph below. The total volume of solutions used in each experiment was 100 cm^3. In each experiment, the difference between the highest temperature reached and the initial temperature was calculated, and was plotted against the volume of H_2O_2 used.



What is the concentration of H_2O_2 and the enthalpy change of the reaction per mole of H_2O_2? Assume specific heat capacity of the solution is 4.2 J g^{-1} K^{-1} and the density of the solution is 1 g cm^{-3}.

	$[H_2O_2]$	Enthalpy change of reaction
A	0.536 mol dm^{-3}	+151 kJ mol^{-1}
B	0.214 mol dm^{-3}	+378 kJ mol^{-1}
C	0.536 mol dm^{-3}	+378 kJ mol^{-1}
D	0.214 mol dm^{-3}	+151 kJ mol^{-1}

45 Enthalpy changes for the following reactions can be determined experimentally:

$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$ $\Delta H_{rxn} = -93$ kJ mol^{-1} (1)

$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l)$ $\Delta H_{rxn} = -242$ kJ mol^{-1} (2)

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$ $\Delta H_{rxn} = -906$ kJ mol^{-1} (3)

What is the enthalpy change of formation of NO(g)?

A + 90 kJ mol^{-1}

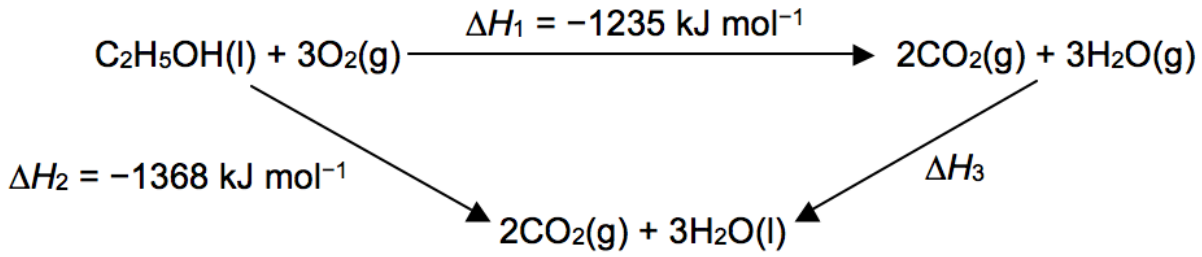
B +180 kJ mol^{-1}

C +360 kJ mol^{-1}

D +1270 kJ mol^{-1}

46 The standard enthalpy change of vapourisation, ΔH_{vap}, is defined as the enthalpy change when one mole of a substance changes from liquid state to the gaseous state under standard conditions.

Given the following energy cycle,



what is the entropy change when 108 g of water vapourises at 373 K?

A −2139 J K^{-1}

B −713 J K^{-1}

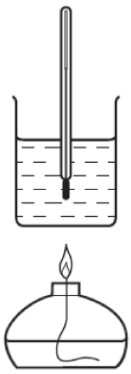
C +713 J K^{-1}

D +2139 J K^{-1}

47 *Use of the Data Booklet is relevant to this question.*

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

The following results were obtained by the student.



initial temperature of the water:	20 ^{o}C
final temperature of the water:	53 ^{o}C
mass of alcohol burner before burning:	259.65 g
mass of alcohol burner after burning:	259.15 g
mass of beaker plus water:	150.00 g
mass of beaker:	50.00 g

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

A 209 J

B 13800 J

C 20700 J

D 22200 J

48 Consider the following enthalpy changes

$Mg(s) \longrightarrow Mg(g)$ ΔH_1 = **P**

$Mg(s) \longrightarrow Mg^{2+}(g) + 2e$ ΔH_2 = **Q**

$Mg(g) \longrightarrow Mg^{+}(g) + e$ ΔH_3 = **R**

$Mg^{+}(g) \longrightarrow Mg^{2+}(g) + e$ ΔH_4 = **S**

$Mg(g) \longrightarrow Mg^{2+}(g) + 2e$ $\Delta H_5 = \mathbf{T}$

Which statements are true?

1 **R** is equal to the 1st IE of Mg.

2 magnitude of (**Q** - **S** – **R**) is greater than that of **P**.

3 **T** is not equal to the sum of **R** and **S**.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

49 The enthalpy change of hydrogenation of ethene is −120 kJ mol^{-1}.

Which of the following are correct?

1 The C-H bond dissociation energy is 60 kJ mol^{-1}.

2 Penta-1,4-diene has an enthalpy change of hydrogenation of −240 kJ mol^{-1}.

3 Ethene has a higher energy content than ethane.

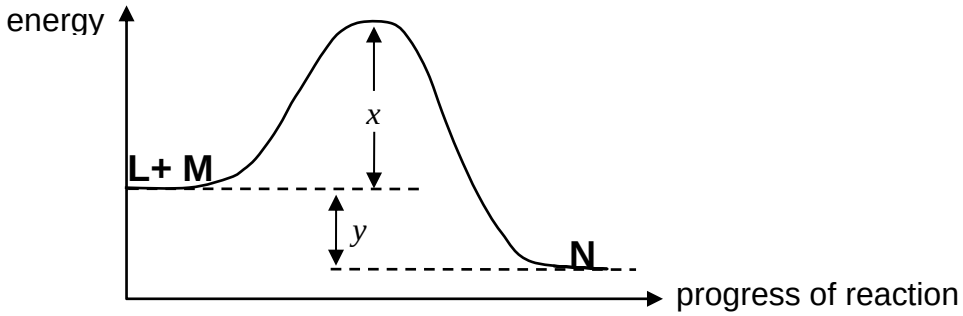
A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

50 The energy profile diagram for a reversible reaction **L+ M** ⇌ **N** is shown below.



Which of the following statements are correct?

1 The activation energy of the backward reaction is $x + y$.

2 The backward reaction is endothermic.

3 The magnitude of the enthalpy change of the backward reaction is $x – y$.

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

51 The conversion of graphite into diamond is a non-spontaneous reaction at 298 K. The entropy change of this reaction is negative at 298 K.

C(graphite) → C(diamond); $\Delta G = +3.08$ kJ mol^{-1}, $\Delta S = –3.3$ J mol^{-1} K^{-1}

Which statements about the enthalpy change for the above reaction are correct?

1 The bond energy of the carbon–carbon bonds in graphite is greater than that in diamond.

2 The enthalpy change of atomisation of diamond is less positive than that of graphite.

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

A 1, 2 and 3

B 1 and 2 only

C 2 and 3 only

D 1 only

52 Which of the following processes are exothermic?

1 $H_2O(l) \rightarrow H_2O(g)$ **2** $Li^+(g) + aq \rightarrow Li^+(aq)$ **3** $O(g) + e^- \rightarrow O^-(g)$

A 1, 2 and 3
B 1 and 2 only
C 2 and 3 only
D 1 only

53 Which of the following processes are more exothermic for gaseous magnesium ions than that of gaseous calcium ions?

1 conversion of gaseous ions into gaseous atoms
2 formation of an ionic oxide lattice
3 hydration of gaseous ions

A 1, 2 and 3
B 1 and 2 only
C 2 and 3 only
D 1 only

Structured Questions

1 (i) When burnt in a limited supply of air, ethane forms carbon monoxide and water. The enthalpy change of the reaction is –1004 kJ per mol of ethane.

Construct a balanced equation that represents this enthalpy change.

(ii) Ethane can be synthesised through the hydrogenation of ethene as shown in the following equation:

$$CH_2{=}CH_2(g) + H_2(g) \longrightarrow CH_3CH_3(g)$$

Calculate the enthalpy change of the hydrogenation reaction, using the following data in addition to the one given in (i):

Enthalpy change of combustion of $CH_2{=}CH_2(g)$ = –1411 kJ mol^{-1}

Enthalpy change of formation of $H_2O(l)$ = –286 kJ mol^{-1}

Enthalpy change of combustion of CO(g) = –283 kJ mol^{-1}

[4]

2 [ACJC 2010] *The use of the Data Booklet is relevant to this question.*

A Group II compound, magnesium chloride, $MgCl_2$, is an important coagulant used in the preparation of soy products. The lattice energy of $MgCl_2$ can be calculated from a Born–Haber cycle using the following data and others found in the *Data Booklet*.

Enthalpy change of atomisation of magnesium	+147 kJ mol^{-1}
Enthalpy change of formation of $MgCl_2$	-641 kJ mol^{-1}
First electron affinity of chlorine	-364 kJ mol^{-1}
Enthalpy change of atomisation of chlorine	+122 kJ mol^{-1}

(i) Explain, with the aid of an equation, what is meant by lattice energy of $MgCl_2$.

(ii) Construct a Born-Haber cycle for the formation of $MgCl_2$ and use it to calculate the lattice energy of $MgCl_2$.

(iii) Explain how you would expect the numerical magnitude of the lattice energy of barium chloride, $BaCl_2(s)$ to compare with that of $MgCl_2(s)$.

(iv) The enthalpy change of formation of a hypothetical ionic solid, magnesium (i) chloride, MgCl, is –125 kJ mol^{-1}.

Using the data given above, calculate the enthalpy change of reaction, $\Delta H_{reaction}$ for the following reaction:

$$2MgCl(s) \rightarrow MgCl_2(s) + Mg(s)$$

Hence comment on the stability of MgCl(s) relative to that of $MgCl_2(s)$.

[Total: 12]

3 [AJC 2010] The Al_2O_3–Cr_2O_3 ceramic is a solid solution of oxides of aluminium and chromium used in incinerators as refractory materials due to the ability to withstand high temperatures.

(a) The ionic radii of Al^{3+} and Cr^{3+} and the enthalpy change of hydration, ΔH_{hyd}, of Al^{3+} are listed below.

	ionic radii / nm	$\otimes H_{hyd}$ / kJ mol^{-1}
Al^{3+}	0.050	–4690
Cr^{3+}	0.069	

How would you expect the ΔH_{hyd} of Cr^{3+} to compare with that of Al^{3+}?

(b) When $AlCl_3$ is dissolved in a large quantity of water, hydrated ions are formed.

$$AlCl_3(s) \xrightarrow{H_{soln}} Al^{3+}(aq) + 3Cl^-(aq)$$

	ΔH / kJ mol^{-1}
standard enthalpy change of atomisation of Al	+326
first electron affinity of Cl	–364
standard enthalpy change of hydration of Cl^-	–381
standard enthalpy change of formation of $AlCl_3$	–704
standard enthalpy change of sublimation of $AlCl_3$	+121

(i) Construct an energy level diagram to calculate the standard enthalpy change of solution, ΔH_{soln}, of $AlCl_3$. Incorporate the data given above as well as relevant data from (**b)** and the *Data Booklet*.

(ii) Using the energy cycle in **(b)(i)** and relevant data from (**b)**, calculate the average bond energy of the Al–Cl bond.

4(a) [CJC 2010] In the process of refining crude oil, shorter and useful hydrocarbons are obtained through catalytic cracking. Silica, SiO_2, a catalyst used in catalytic cracking, can be obtained by the hydrolysis of silicon chloride, $SiCl_4$.

Enthalpy change of atomisation of Si	+338 kJ mol^{-1}
Enthalpy change of formation of $SiCl_4(g)$	-610 kJ mol^{-1}

Using the above information and relevant data from the *Data Booklet*, complete an energy level diagram and hence calculate the average bond energy of Si-Cl*.*

[4]

(b) Through catalytic cracking of decane, a variety of hydrocarbons such as methane, propane and cyclohexene can be obtained.

1.50 g of cyclohexene, C_6H_{10} was burned under a container of 125 g of water causing a temperature rise of the water from 15 ^oC to 75 ^oC. The process was found to be 80% efficient and carbon dioxide is emitted in the process.

(i) Calculate the standard enthalpy change of combustion of cyclohexene, assuming that specific heat capacity of water is 4.20 J g^{-1} K^{-1}.

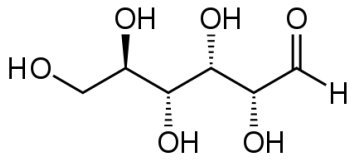
(ii) Suggest a reason why the process is only 80% efficient.

[4]

(c) For safety reasons, the recommended maximum limit for indoor air quality is 600 parts per million by volume(600 ppm) of carbon dioxide. Given that a typical room measures 5 m x 5 m x 3 m, calculate the amount of carbon dioxide gas required to reach the recommended maximum limit at r.t.p.

[2]

5 [CJC 2010] Glucose, $C_6H_{12}O_6$, is a simple sugar used as an energy source in many organisms and a metabolic intermediate. The melting point of glucose is 146 ^{O}C. The structure of glucose is shown below:



An example of the dehydration of a glucose molecule can be seen from the equation below.

$$CH_2OH(CH(OH))_4CHO \xrightarrow{\Delta H_1} CH_2{=}C(OH)\text{-}CH{=}C(OH)\text{-}CH(OH)\text{-}CHO + 2H_2O$$

Glucose **Compound A**

(a) Write an equation to define the standard enthalpy change of formation of glucose. [1]

(b) Using the value(s) given below, calculate ΔH_1.

[2]

Enthalpy change of formation of Glucose	-1271 kJ mol^{-1}
Enthalpy change of formation of **A**	-1085 kJ mol^{-1}
Enthalpy change of combustion of H_2(g)	-286 kJ mol^{-1}
Enthalpy change of atomisation of **C**(s)	+715 kJ mol^{-1}

(c) Use the data given in(b) and any relevant values from the Data Booklet, draw an appropriate energy cycle and hence, find the enthalpy change of sublimation of glucose. [5]

(d) State the total possible number of stereoisomers that **glucose** can display. [1]

6(a) [PJC/08] The decomposition of zinc carbonate is shown below,

$$ZnCO_3(s) \rightarrow ZnO(s) + CO_2(g) \quad \Delta H^{\ominus} = +71 \text{ kJ mol}^{-1}$$

(i) Calculate $\Delta G^{\ominus}$ for the reaction, given its $\Delta S^{\ominus}$ is +174 J mol^{-1} K^{-1} at 298 K and predict whether the reaction is feasible at 298 K.

(ii) Is the decomposition of $ZnCO_3$ feasible at higher temperature? Explain your answer. [4]

(b) Zinc oxide occurs in nature as the mineral zincite. In recent years, it has received considerable attention because of its unique optical, semiconducting, piezoelectric, and magnetic properties.

Using the data provided and relevant data from the *Data Booklet*, construct a labelled Born-Haber cycle to calculate the lattice energy of zinc oxide.

[4]

Standard enthalpy change of formation of zinc oxide / kJ mol^{-1}	-348
Standard enthalpy change of atomisation of zinc / kJ mol^{-1}	+130
Sum of first and second electron affinity for oxygen / kJ mol^{-1}	+657

7 [VJC 2010] Cyclobutane is a colourless gas at room temperature. It decomposes to ethene as shown by the following equation.

(a) To determine ΔH_1, the enthalpy change of the decomposition of cyclobutane to ethene, the enthalpy change of combustion of cyclobutane has to be determined is first by doing the following experiment.

1.00 g of cyclobutane was completely combusted. The heat evolved increased the temperature of 200 g of water by 46.9 ^{o}C. The heat trans.f.er was known to be only 80% efficient. [Specific heat capacity of water = 4.18 J K^{-1} g^{-1}]

(i) Calculate the enthalpy change of combustion of cyclobutane.

(ii) Given the enthalpy change of combustion of ethene is −1422 kJ mol^{-1} and using your answer in **(a)(i)**, calculate ΔH_1. [3]

(b) Ethanol for use as an industrial solvent is made primarily by the acid–catalysed hydration of ethene, represented by the chemical equation below.

$$C_2H_4(g) + H_2O(g) \rightarrow CH_3CH_2OH(l)$$

(i) The enthalpy change of formation of ethanol is –276 kJ mol^{-1}.

With reference to ethanol, define the term *standard enthalpy change of formation*. Include an equation in your answer.

(ii) Using bond energy values from the *Data Booklet*, calculate the enthalpy change for the following reaction.

$$2C(g) + 3H_2(g) + ½O_2(g) \rightarrow CH_3CH_2OH(g)$$

(iii) Using your answer in **(b)(ii)**, the data given in **(b)(i)** and below, draw an energy cycle to calculate the enthalpy change of atomisation of carbon.

$$CH_3CH_2OH(l) \rightarrow CH_3CH_2OH(g) \quad \Delta H_{vap} = +39 \text{ kJ mol}^{-1}$$

[8]

8(a) [AJC/07] Propane and bromine undergo free radical substitution reaction to produce two possible monobrominated propanes.

(i) Outline the reaction mechanism to produce one of the monobrominated products, labelling each step in the mechanism appropriately.

(ii) Draw the structural formula of the other monobrominated product which might be formed.

(iii) Explain why the reaction requires only a flash of ultraviolet light rather than prolonged radiation.

(iv) Different halogens used will result in different rates of free radical substitution of propane. Rate of reaction for different halogens used is as shown in descending order:

$$F_2 > Cl_2 > Br_2 > I_2$$

By quoting suitable *bond energies* from the *Data Booklet*, explain why the *initiation step* for free radical substitution is **not** the rate-determining step.

[8]

(b) One of the possible termination steps in the above free radical substitution reaction is as follows:

$$2 \bullet C_3H_7(g) \longrightarrow C_6H_{14}(g)$$

The standard enthalpy changes of formation, ΔH_f^θ, of $C_3H_7(g)$ and $C_6H_{14}(g)$ are as follows.

	ΔH_f^θ / kJ mol^{-1}
$\bullet C_3H_7$ (g)	+ 103
C_6H_{14} (g)	- 167

(i) Calculate ΔH^θ for this termination step.

(ii) Calculate ΔG^θ for this termination step.(Given $\Delta S^\theta = -190\ J\ K^{-1}\ mol^{-1}$)

(iii) How do the values of ΔH^θ and ΔS^θ relate to the changes that occur at the molecular level for this termination step?

(iv) Show that this termination step is **not feasible** at a higher temperature of 2000 K.

(v) Hence, suggest what would be the effect of this higher temperature on the mechanism of the free radical substitution reaction in **(a)**.

[8]

(c) Liquefied propane gas can be used as cooking gas. The liquefied propane would give propane gas at room temperature and pressure. An experiment was conducted to determine the volume of propane gas required to increase the temperature of 1000 g of water by 75 **°C**

(i) Write a balanced equation for the combustion of propane under standard conditions.

(ii) Calculate the volume of propane gas required for the above experiment.

[heat capacity of water = 4.2 J K^{-1} g^{-1} and $\Delta H^\theta_{c(\text{propane})}$ = – 2220 kJ mol^{-1}]

(iii) State any assumption(s) that you may have made in the calculation above.

[4]

9 [MJC/08] Hydrazine is used as rocket fuel and to prepare gas precursors used in air bags. Approximately 260 thousand tonnes of hydrazine are manufactured annually.

(a) Liquid hydrazine undergoes combustion according to the following equation:

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2H_2O(l)$$

A chemist conducted an experiment to determine the standard enthalpy change of combustion of hydrazine. In the experiment, 0.210 g of hydrazine was burnt as fuel to heat up a beaker containing 200 cm^3 of water. The temperature of water rose by 4 ^oC. You may assume the process has 80% efficiency.

(i) Explain what is meant by standard enthalpy change of combustion of hydrazine. [1]

(ii) Calculate the standard enthalpy change of combustion of hydrazine. [2]

(iii) Given the following data:

enthalpy change of formation of steam = - 242 kJ mol^{-1}

enthalpy change of vapourisation of water = + 44 kJ mol^{-1}

and using your result from **(b)(ii),** draw an appropriate energy cycle to determine the standard enthalpy of formation of hydrazine.

[3]

(b) The standard enthalpy change of formation of hydrazine gas is +235 kJ mol^{-1}.

(i) Using appropriate data from the *Data Booklet*, draw an energy level diagram and calculate the average bond energy of N-H bond in hydrazine.

[3]

(ii) Suggest a reason for the difference in the N-H bond energy value obtained from(**c)(i)** with the value given in the Data Booklet.

[1]

10 [NYJC/08] Many organic compounds exhibit various types of isomerism, e.g. structural isomerism, geometric isomerism and optical isomerism.

(a) The cis-trans inter-conversion of but-2-ene isomers does not occur spontaneously, but it can be made to happen by treating the alkene with a strong acid catalyst.

$$\text{cis-but-2-ene(g)} \rightleftharpoons \text{trans-but-2-ene(g)}$$

(i) Given that the K_p = 3.4 for the above equilibrium, calculate the mole fraction of trans-but-2-ene present at equilibrium.

(ii) The hydrogenation of butene is given by: $C_4H_8(g) + H_2(g) \rightarrow C_4H_{10}(g)$.

However, hydrogenation of the two isomers gives different enthalpy changes as shown:

enthalpy change of hydrogenation of cis-but-2-ene = – 120 kJ mol^{-1}

enthalpy change of hydrogenation of trans-but-2-ene = – 116 kJ mol^{-1}

Calculate the enthalpy change for the cis-trans isomerism of but-2-ene.

(iii) Hence, predict an approximate value of ΔG for the cis-trans isomerism of but-2-ene, showing your reasoning.

[6]

11 In atmospheric pollution, nitrogen oxides commonly refer to a mixture of nitric oxide, NO, and nitrogen dioxide, NO_2, collectively represented as NO_x.

Nitrogen dioxide, which can be classified as a free radical, exists in equilibrium with dinitrogen tetroxide, N_2O_4, a powerful oxidiser which is highly toxic and corrosive.

$$N_2O_4(g) \rightleftharpoons 2NO_2(g)$$

(i) The equilibrium constant K_c of a reaction is related to the absolute temperature, T(in Kelvin), of the reacting conditions via the equation:

$$\lg K_{c,T2} - \lg K_{c,T1} = \frac{\left[\Delta H\left(\frac{1}{T_1} - \frac{1}{T_2}\right)\right]}{2.30R}$$

where a change in temperature from T_1 to T_2 changes the equilibrium constant from $K_{c,T1}$ to $K_{c,T2}$ respectively and R is the molar gas constant.

Given that the equilibrium constant changes from 0.05 mol dm^{-3} at 300 K to 15 mol dm^{-3} at 400 K for the above reaction, calculate the enthalpy change for the reaction. State a major assumption made in this calculation.

(ii) Predict the sign of $\otimes S$ for the reaction: $N_2O_4(g) \rightleftharpoons 2NO_2(g)$, showing your reasoning.

(iii) Using your answers from part **(iii)** and **(iv)**, deduce how the spontaneity of the reaction changes with temperature.

[10]

12 (NYJC 12) Given the following thermochemical data:

Reaction	**ΔH^{θ}/ kJ mol ^{1}**
C(graphite) + $2H_2(g)$ → $CH_4(g)$	75.0
C(graphite) + $O_2(g)$ → $CO_2(g)$	−393.5
$H_2(g)$ + ½$O_2(g)$ → $H_2O(l)$	−285.9

(a) With the aid of an energy cycle, calculate the enthalpy change for the reaction

$CH_4(g) + 2O_2(g) \rightarrow 2H_2O(l) + CO_2(g)$

(b) The experimental enthalpy change is −801.7 kJ mol^{-1} for the following reaction

$$CH_4(g) + 2O_2(g) \rightarrow 2H_2O(g) + CO_2(g) \text{ ……..}\textbf{(I)}$$

(i) Calculate the enthalpy change of vaporisation of water at 298K.

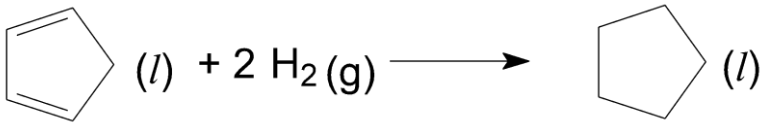
(ii) Using bond energy data from the *Data Booklet*, calculate another value for ΔH for reaction **(I)** in **(b)**. Account for any differences between your answer and the given experimental value in **(b)**.

13 (NJC 13) Cyclopentane, C_5H_{10}, is highly flammable and is commonly used as motor fuel.

(i) A student conducted an experiment to determine the enthalpy change of combustion of cyclopentane. She concluded that 10 grams of liquid cyclopentane is required to bring the temperature of 1 kg of water to its boiling point from its initial temperature of 27 °C.

Given that the process is only 65% efficient in terms of energy transferred, calculate the enthalpy change of combustion of liquid cyclopentane.

(ii) Cyclopentadiene, C_5H_6, can undergo catalytic hydrogenation to form back cyclopentane, C_5H_{10}.



With the use of the following data and your answer to **(i)**, construct an energy cycle and use it to calculate the enthalpy change for the hydrogenation of liquid cyclopentadiene.

	ΔH/ kJ mol^{-1}
Enthalpy change of combustion of gaseous cyclopentadiene	−2800
Enthalpy change of vapourisation of cyclopentadiene	+29
Enthalpy change of combustion of hydrogen gas	−286

14 (HCI 14) Lattice energies can be obtained from the experimental data in a Born-Haber cycle, and they can also be estimated using a pure ionic model represented by the Kapustinskii equation:

$$L.E. = \frac{k\ v\ z_+\ z_-}{r_+ + r_-}$$

where z_+ and z are positive or negative integers giving the charges on the ions, r_+ and r are the ionic radii in nm, v is the number of ions in the formula unit, and k is a constant with a value of 107.9 nm kJ mol^{-1}.
For example, using the Kapustinskii equation, the lattice energy of sodium chloride is calculated as follows:

$$L.E. = \frac{107.9 \times 2 \times (+1) \times (-1)}{(0.095 + 0.181)} = -782\ kJ\ mol^{-1}$$

(i) The table below shows numerical values of lattice energies for sodium chloride and two silver halides. These have been determined from experimental data or theoretically calculated.

compound	experimental value / kJ mol^{-1}	theoretical value / kJ mol^{-1}
NaC*l*	−781	−782
AgF	−967	−824
AgI	−889	−618

Explain the differences between the experimental and theoretical values of lattice energy for AgF and AgI.

(ii) Use the Kapustinskii equation to calculate the lattice energies of calcium(I) chloride and calcium(II) chloride. As calcium(I) chloride is not known to exist, assume that the ionic radius of Ca^+ is similar to that of the neighbouring ion K^+, which is 0.133 nm.

(iii) Use the data below to construct a Born-Haber cycle to calculate the enthalpy change of formation of calcium(I) chloride, CaC*l*. Label each energy level of your cycle, and draw arrows representing the energy terms involved. Use symbols or numbers to represent these energy terms.

enthalpy change of atomisation of calcium = +178 kJ mol^{-1}

electron affinity of chlorine = −349 kJ mol^{-1}

(iv) Construct an energy cycle to calculate the enthalpy change of the following disproportionation reaction:

$$2CaCl(s) \rightarrow Ca(s) + CaCl_2(s)$$

(v) Which of your answers, **(b)(iii)** or **(b)(iv)**, provides support for the argument that calcium(I) chloride does not exist? Explain your reasoning.

1 5 (RV 14) Phosphorus pentachloride can react with phosphorus pentaoxide to produce phosphoryl chloride.

$$6PCl_5(g) + P_4O_{10}(s) \rightarrow 10POCl_3(g)$$

Given the following data, construct an energy cycle and determine the enthalpy change of the reaction above.

$P_4(s) + 6Cl_2(g) \rightarrow 4PCl_3(g)$ $\quad \Delta H_r = -1226$ kJ mol^{-1}

$P_4(s) + 5O_2(g) \rightarrow P_4O_{10}(s)$ $\quad \Delta H_r = -2967$ kJ mol^{-1}

$PCl_3(g) + Cl_2(g) \rightarrow PCl_5(g)$ $\quad \Delta H_r = -84.2$ kJ mol^{-1}

$PCl_3(g) + \frac{1}{2}O_2(g) \rightarrow POCl_3(g)$ $\quad \Delta H_r = -286$ kJ mol^{-1}

16 Potassium hydroxide is formed when oxides of potassium dissolve in water. The following data relate to the energy changes which occur when potassium hydroxide, KOH, and calcium hydroxide, $Ca(OH)_2$, dissolve in water.

	potassium	Calcium
Hydration energy of metal ion / kJ mol^{-1}	–321	–1650
Lattice energy of metal hydroxide / kJ mol^{-1}	–816	–2580
Enthalpy change of solution of metal hydroxide / kJ mol^{-1}	–55	**y**
Extent of hydration of metal ion (Average number of attached water molecules per metal ion)	4	13

(i) Define *standard enthalpy change of solution*.

(ii) Suggest how you would expect the temperature to change when potassium hydroxide is dissolved in water.

(iii) Explain why the hydration energy of Ca^{2+} is more exothermic than that of K^+.

(iv) Draw an energy level diagram to determine the hydration energy of hydroxide ion.

(v) Hence, calculate the value of **y**, the enthalpy change of solution of $Ca(OH)_2$.

(vi) Solubility is also controlled by the entropy change which accompanies dissolution. Use the extent of hydration from the table to compare qualitatively the entropy changes of solution of KOH and $Ca(OH)_2$.

(vii) Suggest and explain one other factor which contributes to the total entropy change when $Ca(OH)_2$ dissolves in water.

[10]

Answers

7) Energetics and Entropy MCQ

1	**D**	**11**	**C**	**21**	**D**	**31**	**A**	**41**	**D**		
2	**B**	**12**	**B**	**22**	**A**	**32**	**C**	**42**	**C**		
3	**C**	**13**	**C**	**23**	**D**	**33**	**D**	**43**	**A**		
4	**A**	**14**	**D**	**24**	**B**	**34**	**B**				
5	**A**	**15**	**B**	**25**	**B**	**35**	**B**	**45**	**A**		
6	**A**	**16**	**A**	**26**	**B**	**36**	**B**	**46**	**C**	**51**	**A**
7	**A**	**17**	**D**	**27**	**A**	**37**	**C**	**47**	**B**	**52**	**C**
8	**A**	**18**	**A**	**28**	**C**	**38**	**C**	**48**	**D**	**53**	**A**
9	**C**	**19**	**D**	**29**	**C**	**39**	**A**	**49**	**C**		
10	**B**	**20**	**A**	**30**	**B**	**40**	**B**	**50**	**B**		