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NZQA | Mana Tohu Mātauranga o Aotearoa
New Zealand Qualifications Authority

Level 3 Chemistry 2025

91390 Demonstrate understanding of thermochemical principles and the properties of particles and substances

Credits: Five

Achievement	Achievement with Merit	Achievement with Excellence
Demonstrate understanding of thermochemical principles and the properties of particles and substances.	Demonstrate in-depth understanding of thermochemical principles and the properties of particles and substances.	Demonstrate comprehensive understanding of thermochemical principles and the properties of particles and substances.

Check that the National Student Number (NSN) on your admission slip is the same as the number at the top of this page.

You should attempt ALL the questions in this booklet.

A periodic table and other reference material are provided in the Resource Booklet L3–CHEMR.

If you need more room for any answer, use the extra space provided at the back of this booklet.

Check that this booklet has pages 2–12 in the correct order and that none of these pages is blank.

Do not write in the margins (⋮⋮⋮). This area will be cut off when the booklet is marked.

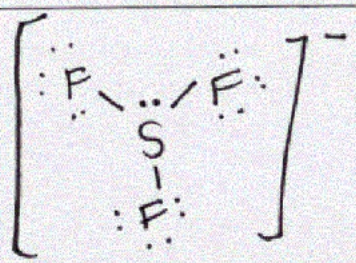
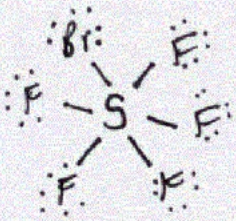
YOU MUST HAND THIS BOOKLET TO THE SUPERVISOR AT THE END OF THE EXAMINATION.

Excellence TOTAL 22

QUESTION ONE

Sulfur forms a range of different fluoride-containing ions and molecules.

(a) Complete the table below:

	SF_3^-	$SBrF_5$
Lewis structure		
Name of shape	Trigonal pyramid	Octahedral

(b) The Lewis structure and shape for sulfur tetrafluoride, SF_4, are shown below.

seesaw

(i) Identify and explain the polarity of SF_4.

In SF_4, there are four S–F polar bonds. Fluorine is the most electronegative atom covalently bonded to sulfur atom. There are four bonding pairs and one lone pair around the central 'S' atom. This gives So, the molecule is asymmetrically arranged. Therefore, SF_4 has a molecular dipole since the effect of bond dipoles do not cancels out and the overall SF_4 molecule is polar.

(ii) Justify why the F–S–F bond angles in SF_4 are different to the F–S–F bond angles in $SBrF_5$.

In SF_4, there are four regions of electron density around the central S atom. These regions repels for maximum separation into trigonal bipyramidal shape to minimise repulsion. There are four bonding regions and one non-bonding region (lone pair) Therefore the shape of SF_4 is 'See-Saw' with a bond angle of 90° and 120°. However, In $SBrF_5$, there are 6 regions of electrons density around the central S atom. These regions repels for maximine separation into octahedral shape to minimise repulsion. There are six bonding regions and no non-bonding regions. Therefore, the overall shape of $SBrF_5$ is octahedral with bond angles of approx 90°.

So, the arrangement of atoms around the central atom and whether there are any lone pairs around the central atom determine the bond angles of the molecules.

(c) Finely powdered sulfur, $S(s)$, readily reacts with fluorine gas, $F_2(g)$, in an exothermic reaction to produce sulfur hexafluoride, $SF_6(g)$. The equation for the reaction is:

$$S(s) + 3F_2(g) \rightarrow SF_6(g)$$

Justify, in terms of the entropy changes of the system and surroundings, why the reaction is spontaneous.

The entropy of the system increases since solid sulfur and gaseous fluorine produced gaseous molecule. Even though there are four moles of the reactants side and one mole of SF_6 on the products side, the produced molecule is gaseous. The solid particles are more ordered than gaseous particles. So, there is a greater dispersal of matter and energy in the system and the particles in the system becomes more disordered. The entropy of the surroundings increases since the reaction is exothermic. The heat energy has been released in the surroundings, the particles in the surroundings gains kinetic / heat energy and starts to move at a faster speed / extent. So, the particles becomes more disordered and there is greater dispersal of matter and energy in the surroundings. (Continued answer on extra page ---)

QUESTION TWO

(a) (i) Complete the table below.

Symbol	Electron configuration (use *s*, *p*, *d* notation)
Cl 17	$1s^2 2s^2 2p^6 3s^2 3p^5$
Cu 29	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^9$
Fe^{3+} 23	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$

(ii) Iron, Fe, can form two different ions, Fe^{2+} and Fe^{3+}. Their ionic radii are given below:

Fe^{2+}	92 pm
Fe^{3+}	79 pm

Explain why the radius of the Fe^{2+} ion is larger than the radius of the Fe^{3+} ion.

Both Fe^{2+} and Fe^{3+} has different number of protons (26). So, the total numbers of electrons occupied shells are also different and there is different shielding / repulsion by the inner electrons to the valence electrons. However, Fe^{3+} has one extra proton than Fe^{2+}. Fe^{3+} has 29 protons and Fe^{2+} has 28 protons, So, Fe^{2+} increasi the valence electrons of Fe^{3+} receives greater electrostatic attraction fora by the positively charged nucleus and valence electrons are pulled with more strength closer to the nucleus decreasing the radius of Fe^{3+} ion and Fe^{2+} ions receives less electrostatic attraction by nucleus as compared to Fe^{3+}. So, Fe^{2+} ion has larger radius than Fe^{3+} ions.

(b) (i) Identify all the types of attractive forces between particles of the following substances in their liquid state.

Substance	$\Delta_{vap}H^{\circ}$/kJ mol^{-1}	Attractive forces
Hydrogen bromide, $HBr(\ell)$	17.3	Temporary dipole, permanent dipole
Hydrogen fluoride, $HF(\ell)$	25.2	Hydrogen bonding, Temporary dipole, permanent dipole
Bromine, $Br_2(\ell)$	29.6	Temporary dipole

H-Br: H-F: Br–Br:

(ii) Explain why the $\Delta_{vap}H°$ of HF is higher than that of HBr.

Both HF and HBr has temporary dipoles and permanent dipoles. Permanent dipole ~~oo~~ due to polar H-F bonds in HF and H-Br polar bonds in HBr. However, HF also has a hydrogen bonding which is the strongest of all three attractive forces because fluorine is the most electronegative atom covalently bonded to hydrogen. So, more amount of heat energy will be required to break all intermolecular forces between HF(l) to convert to HF(g). ~~Even though~~ Even though HBr has a greater electron cloud to greater molar mass than HF. The hydrogen bonding is the strongest of all Attractive forces and require more heat to overcome hydrogen bonding between HF molecules. There $\Delta vapH°$ of HF is higher than $\Delta vapH°$ of HBr.

(iii) Justify why Br_2 has the highest $\Delta_{vap}H°$ of the three substances.

Br_2 ~~has~~ only has the temporary dipole forces. However, Br_2 has the greatest molar mass as compared to both HF and HBr. So, Br_2 has a bigger electron cloud and the strength of the temporary dipoles is the strongest of the temporary dipole forces of HF and HBr. Therefore, a greater amount of heat energy will be required to overcome ~~all~~ the attractive forces in $Br_2(l)$ to convert to gaseous Br_2. Even though, HF and HBr has permanent dipole and ~~for~~ alongside temporary dipole and HF also has the hydrogen bonding, they ~~greater electron cloud of~~ have a very small electron cloud as compared to Br(l). Therefore, Br_2 has the highest $\Delta vapH°$ of all three substances.

(c) Dinitrogen tetroxide, N_2O_4, was reacted with methylhydrazine, CH_3NHNH_2. The equation for the reaction is:

$$4CH_3NHNH_2(\ell) + 5N_2O_4(\ell) \rightarrow 4CO_2(g) + 12H_2O(g) + 9N_2(g)$$

(i) Calculate the standard enthalpy change of this reaction, using the following data.

$C(s) + 3H_2(g) + N_2(g) \rightarrow CH_3NHNH_2(\ell)$	$\Delta_f H^\circ = +54.1$ kJ mol^{-1} = −216.4
$C(s) + O_2(g) \rightarrow CO_2(g)$	$\Delta_f H^\circ = -394$ kJ mol^{-1} = −1576
$5(N_2(g) + 2O_2(g) \rightarrow N_2O_4(\ell)$	$\Delta_f H^\circ = -20.0$ kJ mol^{-1}) = 100
$12(H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$	$\Delta_f H^\circ = -286$ kJ mol^{-1}) = −3432
$H_2O(\ell) \rightarrow H_2O(g)$	$\Delta_f H^\circ = +40.7$ kJ mol^{-1} = 488.4

$4(CH_3NHNH_2(\ell) \longrightarrow N_2(g) + 3H_2(g) + C(s)) = \Delta_f H^\circ = +54.1$ kJ mol^{-1}

$\rightarrow 4CH_3NHNH_2(\ell) \rightarrow 4N_2(g) + 12H_2(g) + 4C(s) = \Delta_f H = -216.4$ kJ mol^{-1}

$4(C(s) + O_2(g) \rightarrow CO_2(g) \cdot \Delta_f H^\circ = -394$ kJ mol^{-1})

$\rightarrow 4C(s) + 4O_2(g) \rightarrow 4CO_2(g) \quad \Delta_f H^\circ = -1576$ kJ mol^{-1}

$\rightarrow 12H_2O(g) \rightarrow 12H_2O(g) \quad \Delta_f H^\circ = +488.4$ kJ mol^{-1}

$\rightarrow 5N_2O_4(\ell) \rightarrow 5N_2(g) + 10O_2(g) \cdot \Delta_f H^\circ = +100$ kJ mol^{-1}

$\rightarrow 12H_2(g) + 6O_2(g) \rightarrow 12H_2O(\ell) \cdot \Delta_f H^\circ = -3432$ kJ mol^{-1}

$4CH_3NHNH_2(\ell) + 5N_2O_4(\ell) \rightarrow 9N_2(g) + 4CO_2(g) + 12H_2O(g)$

$\Delta_f H^\circ = -4636$ kJ mol^{-1}

(ii) Explain the effect on the standard enthalpy change calculated in part (c)(i) if the water was produced as a liquid.

If the water was produced as liquid rather than gas, the reaction would have been more exothermic because in gaseous water all intermolecular forces needs to be broken, so more amount of heat energy is required to all Intermolecular forces between water molecules. So, reaction would have been less exothermic (since heat energy is used) But if in liquid form, not all intermolecular forces needs to be broken. So less amount of heat energy will be required to break intermolecular forces in liquid water than gaseous water. ~~The extra heat energy used is~~ No so, more heat will be released and reaction would be more exothermic.

QUESTION THREE

(a) Nitric acid, HNO_3, is manufactured from ammonia, NH_3. The equation for one of the reactions in the process is:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(\ell) \qquad \Delta_r H^\circ = -1172 \text{ kJ mol}^{-1}$$

(i) Write the balanced equation to represent the standard enthalpy of formation of ammonia in the table below.

Substance	Equation	$\Delta_f H^\circ$ kJ mol^{-1}
$NH_3(g)$	$\frac{1}{2}N_2(g) + 1\frac{1}{2}H_2(g) \rightarrow NH_3(g)$	
$NO(g)$	$\frac{1}{2}N_2(g) + \frac{1}{2}O_2(g) \rightarrow NO(g)$	+90.3
$H_2O(\ell)$	$H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(\ell)$	–286

(ii) Calculate the standard enthalpy of formation for ammonia, $\Delta_f H^\circ(NH_3(g))$, using the data given in part (i).

Q1 $\Delta_f H^\circ$ = [Products] − [Reactants].

$= -1172 = [(90.3 \times 4) + (-286 \times 6)] - [4NH_3]$

$= -1172 = 361.2 - 1716 - 4NH_3$

$= -1172 = -1354.8 - 4NH_3$

$= -1172 + 1354.8 = -4 \cdot NH_3$

$= +182.8 = -4NH_3$

$= \frac{182.8}{4} = -NH_3$

$= -45.7$ KJmol^{-1} is the $\Delta_f H^\circ$ for $NH_3(g)$

(b) (i) Justify, using your knowledge of periodic trends, why magnesium, Mg, has a greater first ionisation energy than strontium, Sr.

Ionisation energy refers to the removal of the first valence electron from a gaseous atom. The ionisation energy increases as you go up the group 2. Ionisation energy depends on the number of valence shell electrons occupied shells thus the shielding from inner electrons and the strength of the electrostatic forces. Mg has three electrons occupied shells whereas Sr has five electrons occupied shells. Even though the number of protons increases down the group, but the no. of electrons shells also increases. The valence electrons of Mg receives less repulsion / shielding from the inner electrons whereas Sr receives a greater shielding from the inner electrons as compared to Mg. Mg receives a greater electrostatic pull by the positively charged nucleus and pulled more strongly towards the nucleus so ~~Mg~~ than Sr due to more number of electron occupied shells. ~~Mg~~ Therefore ~~Mg~~ is be continued on extra page ÷

(ii) The atomic radii of five elements from Period 3 are shown in the table below.

Period 3 element	Na	Mg	Al	Si	P
Atomic radius / nm	0.191	0.160	0.130	0.118	0.110

Explain the trend in atomic radius shown in the table.

The atomic radius ~~decreased~~ increases as you go across the period. As you go across the period 3, all elements have same number of electrons occupied shells (three shells total). So, there is same electron - electron repulsion / shielding by the inner electrons. However, the number of protons increases across the period 3. Na has 11, Mg has 12, Al has 13, Si has 14 and P has 15 protons but number of total electron occupied shell remain the same. So, the valence electrons, as you across the period receives greater electrostatic attraction be and pulled more strongly towards the positively charged

nucleus. So, the distance between nucleus and valence electrons decreases across the period and so does the radius. P has the smallest radii in period 3 as it has the highest no. of protons than Al, Si, Cl, Mg & Na but same electron shell. So, receive highest electrostatic pull by nucleus.

(c) When solid calcium chloride, $CaCl_2(s)$, dissolves in water, the temperature increases from 20.4 °C to 23.2 °C. $\Delta T = 2.8$ °C

$CaCl_2(s) \rightarrow Ca^{2+}(aq) + 2Cl^-(aq)$ $\Delta H = -81.3$ kJ mol^{-1}

Calculate the mass of $CaCl_2(s)$ that must dissolve to cause this temperature increase.

Assume the specific heat capacity of the calcium chloride solution is 4.18 J g^{-1} °C^{-1}.

Assume the mass of the calcium chloride solution is 68.0 g.

$M(CaCl_2) = 111$ g mol^{-1}

$q = mc\Delta T$

$q = 68 \times 4.18 \times 2.8$

$q = -795.872$ J or -0.796 kJ

$\Delta_r H^\circ = \frac{-q}{n} = -81.3 = \frac{0.796}{n}$

$n = \frac{0.796}{81.3} = 0.00978$ moles

$n = \frac{m}{M} = 0.0098 = \frac{m}{111}$

$0.0098 \times 111 = 1.0878 = m$

So, mass of $CaCl_2$ is 1.09 grams

or 1.1 grams (use can say)

Extra space if required.
Write the question number(s) if applicable.

QUESTION NUMBER

1. c) Since, the entropy of system and surrounding both increases, the reaction is spontaneous and favoured. Solid sulfur ($S_{(s)}$) readily reacts with Fluorine gas $F_{2(g)}$. The So, the overall entropy is positive.

2. b) i) Therefore, it requires a greater amount of energy to remove a valence electron from Magnesium's ~~valence shell~~ as it is more strongly pulled by the nucleus and Sr requires less amount of energy to remove one valence shell as it is not that strongly pulled by the nucleus towards itself due to shielding by inner electrons

Extra space if required.
Write the question number(s) if applicable.

QUESTION NUMBER

Extra space if required.
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QUESTION NUMBER

91390

Excellence

Subject: L3 Chemistry

Standard: 91390

Total score: 22

Q	Grade score	Marker commentary
One	M6	Explained the polarity of the molecule by recognising the arrangement of bond dipoles, justified differences in bond angles by relating them to electron geometry, and explained the spontaneity of the process by referring to the change in the surroundings' entropy.
Two	E8	Justified the differences in $\Delta_{vap}H°$ by comparing the strength of the intermolecular forces. Applied Hess's Law to an enthalpy change calculation.
Three	E8	Justified the difference in first ionisation energy and atomic radii in context in terms of the number of energy levels, nuclear charge, repulsion from inner energy levels, and the electrostatic attraction between the nucleus and valence electrons. Calculated the mass of the solid dissolved in a calorimetry calculation.