1. Reaction \( \Delta_r H (\text{kJ mol}^{-1}) \)

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<table>
<thead>
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<tbody>
<tr>
<td>A</td>
<td>(2B(s) + \frac{3}{2}O_2(g) \rightarrow B_2O_3(s))</td>
<td>-1273</td>
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<tr>
<td>B</td>
<td>(B_2H_6(g) + 3O_2(g) \rightarrow B_2O_3(s) + 3H_2O(g))</td>
<td>-2035</td>
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<tr>
<td>C</td>
<td>(H_2(g) + \frac{1}{2}O_2(g) \rightarrow H_2O(l))</td>
<td>-286</td>
</tr>
<tr>
<td>D</td>
<td>(H_2O(l) \rightarrow H_2O(g))</td>
<td>+44</td>
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The enthalpy change for a reaction such as B can be written as the difference between the enthalpy of formation of reactants and products:

\[ \Delta_{\text{rxn}} H^\circ = \Sigma m \Delta f H^\circ (\text{products}) - \Sigma n \Delta f H^\circ (\text{reactants}) \]

For reaction B this becomes:

\[ \Delta_{\text{rxn}} H^\circ = [\Delta f H^\circ (B_2O_3(s)) + 3 \Delta f H^\circ (H_2O(g))] - [\Delta f H^\circ (B_2H_6(g)) + 3 \Delta f H^\circ (O_2(g))] \]

Reaction A corresponds to the heat of formation of \(B_2O_3(s)\) so \(\Delta f H^\circ (B_2O_3(s)) = -1273 \text{ kJ mol}^{-1}\).

- Reaction C corresponds to formation of \(H_2O(l)\) and reaction D corresponds to the energy that must then be supplied to vaporize the liquid into \(H_2O(g)\):
  
  \[ \Delta f H^\circ (H_2O(g)) = (-286 + 44) \text{ kJ mol}^{-1} = -242 \text{ kJ mol}^{-1} \]

- The enthalpy change for reaction B is \(\Delta_{\text{rxn}} H^\circ = -2035 \text{ kJ mol}^{-1}\).
- As \(O_2\) exists as a gas under standard conditions, its enthalpy of formation is zero.

Putting all this together gives:

\[-2035 \text{ kJ mol}^{-1} = ([-1273 + 3 \times -242] - [\Delta f H^\circ (B_2H_6(g)) + 3 \times 0]) \text{ kJ mol}^{-1} \]

Hence,

\[ \Delta f H^\circ (B_2H_6(g)) = +36 \text{ kJ mol}^{-1}. \]

2. As in Q1, the enthalpy change for a reaction can be written as the difference between the enthalpy of formation of reactants and products:

\[ \Delta_{\text{rxn}} H^\circ = \Sigma m \Delta f H^\circ (\text{products}) - \Sigma n \Delta f H^\circ (\text{reactants}) \]

(a) \(2\text{NH}_3(g) + 3\text{O}_2(g) + 2\text{CH}_4(g) \rightarrow 2\text{HCN}(g) + 6\text{H}_2\text{O}(g)\)

As \(O_2\) exists as a gas under standard conditions, its enthalpy of formation is zero. The remaining enthalpies of formation are given.

\[ \Delta_{\text{rxn}} H^\circ = [2 \Delta f H^\circ (\text{HCN}(g)) + 6 \Delta f H^\circ (\text{H}_2\text{O}(g))] - [2 \Delta f H^\circ (\text{NH}_3(g)) + 2 \Delta f H^\circ (\text{CH}_4(g))] \]

\[ = ([2 \times 135 + 6 \times -242] - [2 \times -46 + 2 \times -75]) \text{ kJ mol}^{-1} = -940 \text{ kJ mol}^{-1} \]
(b) \[ \text{NH}_3(g) + \text{HCl}(g) \rightarrow \text{NH}_4\text{Cl}(s) \]

\[ \Delta_{\text{rxn}}H^\circ = [\Delta_{f}H^\circ(\text{NH}_4\text{Cl}(s))] - [\Delta_{f}H^\circ(\text{NH}_3(g)) + \Delta_{f}H^\circ(\text{HCl}(g))] \]

\[ = ([-314] - [-46 + -92]) \text{ kJ mol}^{-1} = -176 \text{ kJ mol}^{-1} \]

3. (a) A general exothermic reaction:

(b) A general endothermic reaction:

4. Water vapour

\[ \text{water} \]

\[ \text{ice} \]
5. H-bonding occurs between the electronegative atom (F, O and N) on one molecule and the electropositive hydrogen atom on a second molecule. The electronegativities decrease in the order F > O > N so that the H-F bond is the most polar with the most positive H atom and the N-H bond is the least polar with the least positive H atom. The strength of the H-bonding therefore decreases in the order HF > H$_2$O > NH$_3$.

6. If the two atoms involved in a bond are different, the bond will be polar and will possess a dipole. Unless the shape of the molecule is so symmetrical that the dipoles on each bond cancel, the overall molecule will have a dipole moment.

(a) Cl$_2$O

Oxygen is more electronegative than chlorine so it has a small negative charge and chlorine has a small positive charge.

There are two bonds and two lone pairs on oxygen so the molecule is bent leading to an overall dipole moment as shown.

(b) CH$_3$CH$_2$Cl

Chlorine is the more electronegative than carbon so it has a small negative charge and the carbon it is bonded to has a small positive charge. The molecule has a dipole moment as shown.

The difference in electronegativity between C and H is small - the C-H bonds are slightly polar.

(c) CH$_3$NH$_2$

Nitrogen is more electronegative than carbon and hydrogen so it has a small negative charge and the carbon and hydrogen atoms it is attached to have small positive charges.

7. Dispersion forces increase with the number of electrons present. They therefore increase as a group is descended and as the molecular size grows.

(a) HCl or HI

Iodine possesses more electrons than chlorine and these are located further from the nucleus and so its electron cloud is more easily distorted.

(b) CH$_3$CH$_3$ or CH$_3$CH$_2$CH$_2$CH$_2$CH$_3$CH$_2$CH$_3$

The longer chain leads to more bonds (more electrons).

(c) H$_2$O or H$_2$Se

Selenium possesses more electrons than oxygen and these are located further from the nucleus and so its electron cloud is more easily distorted.

8. Dipole-dipole forces occur between dipoles and are strongest when these are large.

(a) CHCl$_3$ or CCl$_4$

The C-Cl bond is polar. However, in CCl$_4$, the tetrahedral shape leads to no overall dipole moment as there are negatively charged chlorine atoms on each
corner. In CHCl₃, however, there is an overall dipole moment with negatively charged chlorine and positively charged hydrogen.

(b) CH₃CH₂F or CH₃CH₂Br
Fluorine is more electronegative than bromine so the dipole moment is larger in CH₃CH₂F.

(c) CH₃COCH₃ or CH₃CH₂CH₃
The electronegative oxygen atom leads to a large dipole moment in CH₃COCH₃. As carbon and hydrogen have very similar electronegativities, the C-H bonds in CH₃CH₂CH₃ are not very polar and it has a very small dipole moment and, hence, weak dipole-dipole forces.

9. (i) Dispersion forces exist between all molecules.
(ii) Dipole-dipole forces only exist when dipoles are present.
(iii) Ion-dipole forces only exist when ions and molecules with dipole moments are present.
(iv) Dipole-induced dipole forces require the presence of molecules with dipole moments.
(v) H-bonding only occurs when H is bonded to an electronegative element – it is an especially strong version dipole-dipole interaction.

(a) methane, CH₄
No ions or dipole. H is not bonded to an electronegative element.
Only dispersion forces.

(b) acetone, CH₃COCH₃
The polar C=O bond leads to a dipole moment. No ions are present. H is not bonded to an electronegative element.
Dispersion, dipole-dipole and dipole-induced dipole forces.

(c) ammonia, NH₃
H is bonded to an electronegative element.
H-bonding and dispersion forces.

(d) acetic acid, CH₃COOH
H is bonded to an electronegative element (O-H bonds).
H-bonding and dispersion forces.

10. High boiling points occur when intermolecular forces are strong. Dispersion forces tend to be weaker than dipole-dipole interactions, unless the dipoles are very small. H-bonding tends to be the strongest force, if present.

(a) CH₂Cl₂ or CH₂Br₂
Chlorine is more electronegative than bromine so the dipole moment in CH₂Cl₂ is larger than that in CH₂Br₂ leading to stronger dipole-dipole forces.

(b) HF or HBr
As fluorine is more electronegative than bromine, the H-F bond is very polar and H-bonding, which requires H bonded to an electronegative element, is strong in HF.
(c) CH₃CH₂CH₃ or CH₃CH₂CH₂CH₃CH₃
Dispersion forces are the only intermolecular forces in these compounds. The longer chain has more electrons (more bonds) and so it possesses the stronger dispersion forces.

(d) CH₃CH₂CH₂OH or CH₃CH₂CHO
Both molecules possess dipole-dipole interactions due to the presence of electronegative oxygen, CH₃CH₂CH₂OH, however, contains hydrogen bonded to an electronegative atom so H-bonding is possible.

11. (a) C₂H₆ or C₄H₁₀
Only dispersion forces are present and these are largest in the longer chain alkane as it has more electrons (more bonds).

(b) CH₃CH₂OH or CH₃CH₂F
Both molecules possess dipole moments but CH₃CH₂OH contains hydrogen bonded to an electronegative element so H-bonding is possible.

(c) CH₃COOH or (CH₃)₂CO
Both molecules possess dipole moments but CH₃COOH contains hydrogen bonded to an electronegative element so H-bonding is possible.

12. V (+5): HNO₃, NO₃⁻

HNO₃ O has O.N. = -2. H is attached to a non-metal so O.N. = +1
N must have O.N. = +5 so that overall charge is zero: (+1) + (+5) + (3×-2) = 0

NO₃⁻ O has O.N. = -2. Ion has charge of -1.
N must have O.N. = +5 so that overall charge is -1: (+5) + (3×-2) = -1

IV (+4): NO₂, N₂O₄

NO₂ O has O.N. = -2.
N must have O.N. = +4 so that overall charge is zero: (+4) + (2×-2) = 0

N₂O₄ O has O.N. = -2.
N must have O.N. = +4 so that overall charge is zero: (2×+4) + (2×-2) = 0

III (+3): HNO₂, NO₂⁻

HNO₂ O has O.N. = -2. H is attached to a non-metal so O.N. = +1
N must have O.N. = +3 so that overall charge is zero: (+1) + (+3) + (2×-2) = 0

NO₂⁻ O has O.N. = -2. Ion has charge of -1.
N must have O.N. = +3 so that overall charge is -1: (+3) + (2×-2) = -1

II (+2): NO

NO O has O.N. = -2.
N must have O.N. = +2 so that overall charge is zero: (+2) + (-2) = 0
I (+1): N₂O,

N₂O  O has O.N. = -2.
    N must have O.N. = +1 so that overall charge is zero:  \(2\times+1) + (-2) = 0\)

0: N₂

N₂  N must have O.N. = 0 as its in elemental form) = 0

-I (-1): NH₂OH,

NH₂OH
  O has O.N. = -2. Each H is attached to a non-metal so O.N. = +1
  N must have O.N. = -1 so that overall charge is zero:  \((-1) + (2\times+1) + (-2) + (+1) = 0\)

-H (-2): N₂H₄,

N₂H₄  H is attached to a non-metal so O.N. = +1
    N must have O.N. = -2 so that overall charge is zero:  \((2\times-2) + (4\times+1) = 0\)

-III (-3): NH₃, NH₄⁺

NH₃  H is attached to a non-metal so O.N. = +1
    N must have O.N. = -3 so that overall charge is zero:  \((-3) + (3\times+1) = 0\)

NH₄⁺  H is attached to a non-metal so O.N. = +1. Ion has charge of -1.
    N must have O.N. = -3 so that overall charge is +1:  \((-3) + (4\times+1) = +1\)

13. \(V = 1.0 \times 10^{-3} \text{ mL} = 1.0 \times 10^{-6} \text{ L}\).

As 1 atm = 101.3 kPa,

\[P = 1.3 \times 10^{-8} \text{ kPa} = \frac{1.3 \times 10^{-8}}{101.3} \text{ atm} = 1.3 \times 10^{-10} \text{ atm}\]

With the volume in litres and the pressure in atmospheres, \(R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}\) will be used.

The ideal gas law gives \(PV = nRT\) so the number of moles present is given by:

\[n = \frac{PV}{RT} = \frac{(1.3 \times 10^{-10} \text{ atm})(1.0 \times 10^{-6} \text{ L})}{(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(298 \text{ K})} = 5.2 \times 10^{-18} \text{ mol}\]

1 mol corresponds to \(6.022 \times 10^{23}\) molecules so the number of molecules present is:

\[\text{number of molecules} = (5.2 \times 10^{-18}) \times (6.022 \times 10^{23}) = 3.2 \times 10^6 \text{ molecules (i.e. 3.2 million)}\]

14. The decomposition reaction is:

\[\text{NH}_4\text{NO}_2(s) \rightarrow \text{N}_2(g) + 2\text{H}_2\text{O}(l)\]
The molar mass of NH₄NO₂ is:

$$\text{molar mass} = (14.00 \text{ (N)} + (4 \times 1.008 \text{ (H)}) + 14.00 \text{ (N)} + (2 \times 16.00 \text{ (O)})) \text{ g mol}^{-1}$$

$$= 64.032 \text{ g mol}^{-1}$$

The number of moles of NH₄NO₂ is therefore:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{1.00 \text{ g}}{64.032 \text{ g mol}^{-1}} = 0.0157 \text{ mol}$$

From the decomposition equation, one mole of N₂(g) is produced from every mole of NH₄NO₃. The number of moles of N₂ produced is therefore 0.0125 mol.

With the pressure in atmospheres, $$R = 0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1}$$ will be used giving the volume in litres. $$T = 250 \degree \text{C} = (250 + 273) \text{ K} = 523 \text{ K}.$$

The ideal gas law gives $$PV = nRT$$ so the volume is given by:

$$V = \frac{nRT}{P} = \frac{(0.0157 \text{ mol})(0.08206 \text{ L atm K}^{-1} \text{ mol}^{-1})(523 \text{ K})}{(1.00 \text{ atm})} = 0.678 \text{ L}$$

The smallest number of significant figures is the volume (3) so the answer is also reported to three significant figures.