	Marks
• Write balanced equations for the following nuclear reactions.	4

Nickel-63 undergoes beta decay to become a stable nuclide.	$^{63}_{28}\text{Ni} \rightarrow {}^{63}_{29}\text{Cu} + {}^{0}_{-1}\beta$
An alpha particle is produced from the decay of radon-222.	$^{222}_{86}\text{Rn} \rightarrow {}^{218}_{84}\text{Po} + {}^{4}_{2}\text{He}$
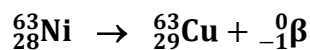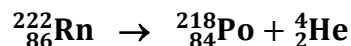
• Direct damage to the DNA of skin cells can be brought about by exposure to ultraviolet radiation of wavelength 300.0 nm. What are the frequency and energy of this radiation?   **2**

> **The wavelength, λ, is related to the energy and the frequency, ν, by the equations:**
>
> $$\nu = \frac{c}{\lambda} \text{ and } E = h\nu = \frac{hc}{\lambda}$$
>
> **Therefore with $\lambda = 300.0$ nm $= 3.000 \times 10^{-7}$ m:**
>
> $$\nu = \frac{(2.998\times10^{8} \text{ m s}^{-1})}{(3.000\times10^{-7} \text{ m})} = 9.993 \times 10^{14} \text{ s}^{-1}$$
>
> $$E = \frac{(6.626\times10^{-34} \text{ J s})\times(2.998\times10^{8} \text{ m s}^{-1})}{(3.000\times10^{-7} \text{ m})} = 6.622 \times 10^{-19} \text{ J}$$
>
> **(As the wavelength is given to four significant figures, this limits the accuracy of the answers to also being four significant figures).**

Frequency: **9.993×10^{14} s^{-1}**	Energy: **6.622×10^{-19} J**

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY

- Complete the following table.

Marks
9

Molecular formula	NH_3	PCl_5	BrF_3
Name	**ammonia**	**phosphorus pentachloride**	**bromine trifluoride**
Lewis structure			
Number of bonding electron pairs on central atom	**3**	**5**	**3**
Number of non-bonding electron pairs on central atom	**1**	**0**	**2**
Molecular shape	**trigonal pyramidal**	**trigonal bipyramdial**	**'T-shaped'**

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|	Marks
What is the ground state electron configuration for the chlorine atom?	2

$1s^2\, 2s^2\, 2p^6\, 3s^2\, 3p^5$ or $[Ne]\, 3s^2\, 3p^5$

- Briefly explain the concept of resonance. Give at least one example.   2

If two or more Lewis structures are equally valid for a molecule, the true structure is none of the structures that is drawn, but an average made up of all these resonance restructures.

In benzene, for example, either of two Lewis structures below is equally valid:
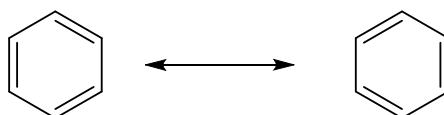
They differ only in are the double and which are the single bonds. The true structure is an average of the two and the ring is a symmetrical hexagon.
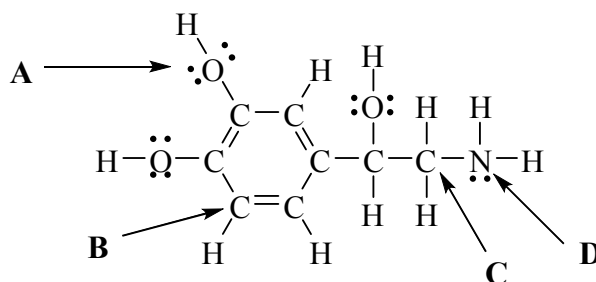
Similarly, the Lewis structure of the carbonate ion, CO_3^{2-}, can be drawn in three ways depending on which oxygen atoms have the negative charges:
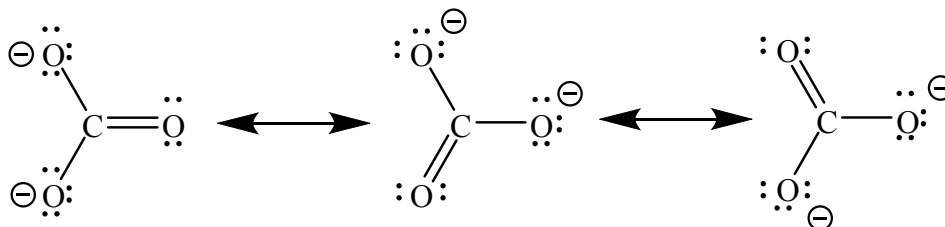
Again, all are equally valid and the real distribution of the electrons is an average of these resonance structures.

- The structure of adrenaline is shown below.   3

Give the approximate bond angles at the indicated atoms.

A: 109.5°	B: 120°	C: 109.5°	D: 109.5°

ANSWER CONTINUES ON THE NEXT PAGE

Which, if any, of the indicated atoms has at least one lone pair of electrons?

An oxygen atom like (A) which is making two bonds has two lone pairs.

A nitrogen atom like (D) which is making three bonds has one lone pair.

Carbon atoms like (B) and (C) which are making four bonds have no lone pairs.

Marks
4

What mass of oxygen is required for the complete combustion of 5.8 g of butane, C_4H_{10}. How many moles of CO_2 and H_2O are produced?
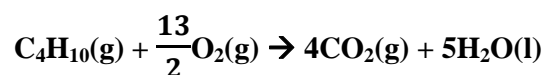
The molar mass of butane, C_4H_{10}, is:

   molar mass = M = (4 × 12.01 (C) + 10 × 1.008 (H)) g mol^{-1} = 58.12 g mol^{-1}

Hence, the 5.8 g corresponds to:

   number of moles = $\dfrac{\text{mass}}{\text{number of moles}} = \dfrac{m}{M} = \dfrac{5.8\ g}{58.12\ g\ mol^{-1}} = 0.10\ mol$

The balanced equation for the combustion of butane is:

   $C_4H_{10}(g) + \dfrac{13}{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(l)$

Hence, 13/2 moles of O_2 are required and 4 moles of CO_2 and 5 moles of H_2O are produced for every mole of C_4H_{10} which combusts. As 0.10 mol of C_4H_{10} is present:

   number of moles of O_2 needed = $\dfrac{13}{2}$ × 0.10 mol = 0.65 mol

   number of moles of CO_2 produced = 4 × 0.10 mol = 0.40 mol

   number of moles of H_2O produced = 5 × 0. 10 mol = 0.50 mol

O_2 has a molar mass of (2 × 16.00) g mol^{-1} = 32.00 g mol^{-1}. Hence the mass of O_2 required is:

   mass = number of moles × molar mass
   = nM = (0.65 mol) × (32.00 g mol^{-1}) = 21 g

2

• A white powder used in paints, enamels and ceramics has the following mass percentage: 69.6% Ba; 6.09% C; 24.3% O. What is its empirical formula?

	Ba	**C**	**O**
amount in 100 g	69.6	6.09	24.3
ratio (divide by atomic mass)	$\dfrac{69.6}{137.34} = 0.507$	$\dfrac{6.09}{12.01} = 0.507$	$\dfrac{24.3}{16.00} = 1.52$
divide by smallest	$\dfrac{0.507}{0.507} \sim 1$	$\dfrac{0.507}{0.507} \sim 1$	$\dfrac{1.52}{0.507} \sim 3$

The simplest possible ratio of Ba:C:O is thus 1:1:3 and the empirical formula is $BaCO_3$.

Answer: **$BaCO_3$**

	Marks
• Lead(II) iodide precipitates when 0.080 M lead(II) nitrate solution (150.0 mL) is added to 0.080 M potassium iodide solution (50.0 mL). Write a balanced ionic equation for the reaction that occurs.	**6**

$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$
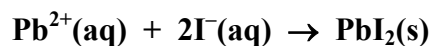
What amount (in mol) of lead(II) iodide precipitates?

Before precipitation, the number of moles of $Pb^{2+}(aq)$ and $I^-(aq)$ present are:

number of moles of Pb^{2+} = concentration × volume

$$= cV = (0.080 \text{ mol L}^{-1}) \times (0.1500 \text{ L}) = 0.012 \text{ mol}$$

number of moles of I^- = (0.080 mol L^{-1}) × (0.0500 L) = 0.0040 mol

The ionic equation shows that 2 moles of I^- are required for every one mole of Pb^{2+}. As there is less I^- present than Pb^{2+}, iodide is the limiting reagent and some of the lead(II) ions are left in solution after precipitation.

One mole of $PbI_2(s)$ is formed for every two moles of $I^-(aq)$ present and hence:

number of moles of $PbI_2(s) = ½ × 0.0040 \text{ mol} = 0.0020 \text{ mol}$

Answer: **0.0020 mol**

What amount (in mol) of $Pb^{2+}(aq)$ ions remain in solution after the reaction?

From above, 0.012 mol of $Pb^{2+}(aq)$ is initially present. The ionic equation shows that one mole of $Pb^{2+}(aq)$ is lost for every one mole of $PbI_2(s)$ formed. As 0.0020 mol of $PbI_2(s)$ precipitates,

number of moles of $Pb^{2+}(aq)$ left = (0.012 – 0.0020) mol = 0.010 mol

Answer: **0.010 mol**

ANSWER CONTINUES ON THE NEXT PAGE

What is the final concentration of NO_3^-(aq) ions remaining in solution after the reaction?

The nitrate is not involved in the reaction so the amount of it is unchanged.

$Pb(NO_3)_2$ contains two moles of NO_3^{2-} for every mole of Pb^{2+}. From above, 0.012 mol of Pb^{2+} is present so there must be (2×0.012) mol = 0.024 mol of NO_3^{2-}(aq).

When the solutions are mixed, the total volume becomes $(150.0 + 50.0)$ mL = 200.0 mL. Hence, the concentration of NO_3^{2-}(aq) becomes:

$$\text{concentration} = \frac{\text{number of moles}}{\text{volume}} = \frac{n}{V} = \frac{0.024 \text{ mol}}{0.2000 \text{ L}} = 0.12 \text{ mol L}^{-1}$$
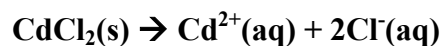
Answer: **0.12 mol L^{-1} = 0.12 M**

	Marks
	4

- Cadmium chloride and cadmium sulfate are both soluble in water. Cadmium carbonate and cadmium hydroxide are both insoluble. Describe, using equations where appropriate, how to convert cadmium chloride into cadmium sulfate.

One possible method is:

(1) Dissolve cadmium chloride in water:

$$CdCl_2(s) \rightarrow Cd^{2+}(aq) + 2Cl^-(aq)$$

(2) Add aqueous sodium hydroxide. This leads to precipitation of cadmium hydroxide which can then be separated from the $Na^+(aq)$ and $Cl^-(aq)$ by filtration.

$$Cd^{2+}(aq) + OH^-(aq) \rightarrow Cd(OH)_2(s)$$
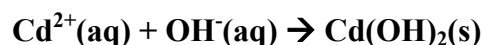
(3) Add the filtered solid to aqueous sulfuric acid. This leads to dissolution as cadmium sulfate is soluble.

$$Cd(OH)_2(s) + 2H^+(aq) \rightarrow Cd^{2+}(aq) + 2H_2O(l)$$
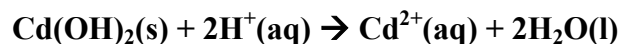
(4) This solution contains only $Cd^{2+}(aq)$, $SO_4^{2-}(aq)$ and $H_2O(l)$ as all of the Cl^- and $Na^+(aq)$ were removed in the filtration in step (2). Simple evaporation of the solvent (water) then leaves $CdSO_4(s)$.

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY

**Marks
6**

- Ammonia, NH_3, is produced from nitrogen and hydrogen gas at high temperatures using the Haber process. At a temperature of 670 K and 50.0 MPa pressure, an equilibrium mixture was found to contain 0.925 mol nitrogen, 2.775 mol hydrogen and 1.50 mol ammonia. Write a balanced equation for the Haber process.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$
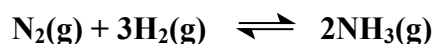
Calculate the mole fraction of each gas in the mixture.

Total number of moles = (0.925 (N_2) + 2.775 (H_2) + 1.50 (NH_3)) mol = 5.20 mol

As mole fraction of A $= X_A = \dfrac{\textbf{number of moles of A}}{\textbf{total number of moles}}$,

$$X_{N_2} = \frac{0.925 \text{ mol}}{5.20 \text{ mol}} = 0.178$$

$$X_{H_2} = \frac{2.775 \text{ mol}}{5.20 \text{ mol}} = 0.534$$

$$X_{NH_3} = \frac{1.50 \text{ mol}}{5.20 \text{ mol}} = 0.288$$

Calculate the partial pressure of each gas.

The mole fraction, X_A, and the partial pressure, p_A, are related by:

$$p_A = X_A \, P_{total}$$

As P_{total} = 50.0 MPa,

$$p_{N_2} = 0.178 \times 50.0 \text{ MPa} = 8.90 \text{ MPa}$$

$$p_{H_2} = 0.534 \times 50.0 \text{ MPa} = 26.7 \text{ MPa}$$

$$p_{NH_3} = 0.288 \times 50.0 \text{ MPa} = 14.4 \text{ MPa}$$

Calculate the value for K_p for the reaction at this temperature.

The equilibrium constant in terms of partial pressures is given by:

$$K_p = \frac{(p_{NH_3})^2}{(p_{N_2})(p_{H_2})^3}$$

The partial pressures need to be in atmospheres to be used in this equation.

ANSWER CONTINUES ON THE NEXT PAGE

The equilibrium constant in terms of partial pressures is given by:

$$K_p = \frac{(p_{NH_3})^2}{(p_{N_2})(p_{H_2})^3}$$

The partial pressures need to be in atmospheres to be used in this equation.

As 1 atm = 101.3 kPa = 1.013×10^5 Pa,

$$p_{N_2} = 8.90 \text{ MPa} = \frac{8.90 \times 10^6}{1.013 \times 10^5} \text{ atm} = 88.0 \text{ atm}$$

$$p_{H_2} = 26.7 \text{ MPa} = \frac{27.2 \times 10^6}{1.013 \times 10^5} \text{ atm} = 263 \text{ atm}$$

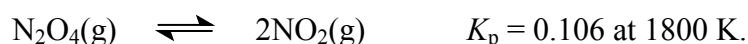$$p_{NH_3} = 14.4 \text{ MPa} = \frac{14.4 \times 10^6}{1.013 \times 10^5} \text{ atm} = 142 \text{ atm}$$

Hence,

$$K_p = \frac{(p_{NH_3})^2}{(p_{N_2})(p_{H_2})^3} = \frac{(142)^2}{(88.0)(263)^3} = 1.26 \times 10^{-5} \quad \text{(dimensionless)}$$

Answer: **1.26×10^{-5}**

	Marks
• The dissociation of gaseous N_2O_4 to NO_2 in the upper atmosphere occurs according to the following equation.	**2**

$$N_2O_4(g) \rightleftharpoons 2NO_2(g) \qquad K_p = 0.106 \text{ at 1800 K.}$$
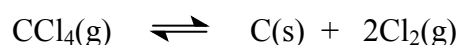
What is the free energy change (in kJ mol^{-1}) for this reaction?

As $\Delta_r G^\circ = -RT\ln K_p$,

$\Delta_r G^\circ = -(8.314 \text{ J K}^{-1} \text{ mol}^{-1}) \times (1800 \text{ K}) \times \ln(0.106)$
$= 33600 \text{ J mol}^{-1} = 33.6 \text{ kJ mol}^{-1}$

Answer: **33.6 kJ mol^{-1}**

• A sample of 0.62 mol CCl_4 was placed in a 2.0 L container and heated to a certain temperature. At equilibrium, $[Cl_2] = 0.060$ M. What is the value of the equilibrium constant K_c for the following reaction at that temperature?	**3**

$$CCl_4(g) \rightleftharpoons C(s) + 2Cl_2(g)$$

The initial concentration of CCl_4 is:

$$\text{concentration} = \frac{\textbf{number of moles}}{\textbf{volume}} = \frac{\textbf{0.62 mol}}{\textbf{2.0 L}} = 0.31 \text{ mol L}^1 = 0.31 \text{ M}$$

The reaction table is:

concentration	CCl$_4$(g)	$\rightleftharpoons$	C(s)	2Cl$_2$(g)
start	**0.31**		**-**	**0**
change	**-x**		**-**	**+2x**
equilibrium	**0.31-x**		**-**	**2x**

As C(s) is a solid, its concentration is essentially constant and is not included in the equilibrium calculations.

At equilibrium, $[Cl_2(g)] = 0.060$ M and hence $x = 0.030$ M. Hence at equilibrium, $[CCl_4(g)] = (0.31 - x)$ M $= 0.28$ M.

The equilibrium constant in terms of the concentrations, K_c, is thus:

$$K_c = \frac{[Cl_2(g)]^2}{[CCl_4(g)]} = \frac{(0.060)^2}{(0.28)} = 0.013$$

Answer: **0.013**

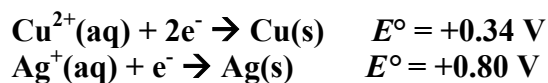**Marks**

2

- A galvanic cell consists of a Cu^{2+}(aq)/Cu(s) and a Ag^+(aq)/Ag(s) half cell. If the voltage of the cell is 0.35 V and the concentration of Cu^{2+}(aq) is 3.5 M, what is the concentration of Ag^+(aq)?

The reduction potentials for the two half cells are (from the data sheet):

$$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s) \quad E° = +0.34 \text{ V}$$
$$Ag^+(aq) + e^- \rightarrow Ag(s) \quad E° = +0.80 \text{ V}$$

The Cu^{2+}(aq) / Cu(s) cell has the lower electrode potential and it is the one that is reversed. Hence, the cell reaction and the standard cell potential are

$$2Ag^+(aq) + Cu(s) \rightarrow 2Ag(s) + Cu^{2+}(aq) \quad E° = (+0.80 \text{ V}) + (-0.34 \text{ V}) = 0.46 \text{ V}$$
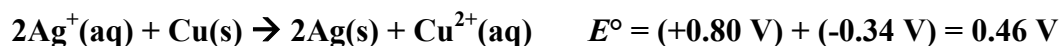
At non-standard concentrations, the electrode potential is given by the Nernst equation:

$$E = E° - \frac{RT}{nF} \ln Q$$

For this two electron reduction, $n = 2$ and $Q = \dfrac{[Cu^{2+}(aq)]}{[Ag^+(aq)]^2}$

When $[Cu^{2+}(aq)] = 3.5$ M, $E = 0.35$ V. Hence, at T = 298 K:

$$(0.35 \text{ V}) = (0.46 \text{ V}) - \frac{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})}{(2 \times 96485 \text{ C mol}^{-1})} \ln\left(\frac{(3.5)}{[Ag^+(aq)]^2}\right)$$

Hence,

$$\ln\left(\frac{(3.5)}{[Ag^+(aq)]^2}\right) = ((0.35 - 0.46) \text{ V}) \times \frac{-(2 \times 96485 \text{ C mol}^{-1})}{(8.314 \text{ J K}^{-1} \text{ mol}^{-1})(298 \text{ K})} = 8.57$$

$[Ag^+(aq)] = 0.026$ M

Answer: **0.026 M**

- How many minutes will be required for a 1.50 A current to electroplate 1.97 g of gold from a solution containing $AuCl_4^-$ ions?

2

1.97 g of gold corresponds to:

$$\text{amount of gold} = \frac{\text{mass}}{\text{atomic mass}} = \frac{1.97 \text{ g}}{196.97 \text{ g mol}^{-1}} = 0.0100 \text{ mol}$$

The oxidation number of gold is +3 (as $AuCl_4^- = Au^{3+} \, 4Cl^-$). Hence, 3 moles of electrons are required to reduce 1 mole of $AuCl_4^-$ to gold. Thus, 0.0300 mol of electrons are required to reduce 0.0100 mol.

ANSWER CONTINUES ON THE NEXT PAGE

1.97 g of gold corresponds to:

$$\text{amount of gold} = \frac{\text{mass}}{\text{atomic mass}} = \frac{1.97 \text{ g}}{196.97 \text{ g mol}^{-1}} = 0.0100 \text{ mol}$$

The oxidation number of gold is +3 (as $AuCl_4^- = Au^{3+}$ $4Cl^-$). Hence, 3 moles of electrons are required to reduce 1 mole of $AuCl_4^-$ to gold. Thus, 0.0300 mol of electrons are required to reduce 0.0100 mol.
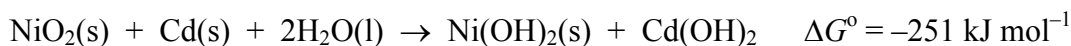
As,

$$\text{moles of electrons} = \frac{It}{F}$$

The time, t, required with current I = 1.50 A to deliver 0.0300 mol is:

$$t = \frac{(0.0300 \text{ mol}) \times (96485 \text{ C mol}^{-1})}{1.50 \text{ A}} = 1930 \text{ s} = 32.2 \text{ minutes}$$

Answer: **32.2 minutes**

- A solar powered light uses a nickel-cadmium battery to store electricity. Calculate the standard voltage for the battery from the following:

$$NiO_2(s) + Cd(s) + 2H_2O(l) \rightarrow Ni(OH)_2(s) + Cd(OH)_2 \quad \Delta G^o = -251 \text{ kJ mol}^{-1}$$

3

The reaction is a two electron process (as Cd is being oxidized to Cd(II) and Ni is being reduced from Ni(IV) to Ni(II)). Hence, n = 2.

The standard electrode potential is linked to Gibbs free energy change by,
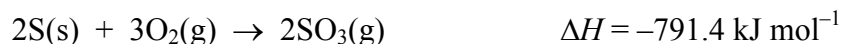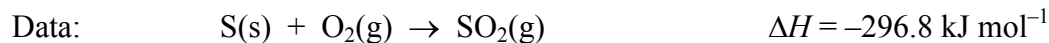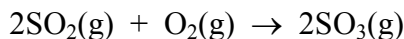
$$\Delta G° = -nFE°$$

Hence with $\Delta G°$ = -251 kJ mol^{-1}

$$-251 \times 10^3 \text{ J mol}^{-1} = -2 \times (96485 \text{ C mol}^{-1}) \times E°$$

$$E° = 1.30 \text{ V}$$

Answer: **1.30 V**

**Marks
2**

- The conversion of SO_2 to SO_3 can occur in the catalytic converters of cars using gasoline containing traces of sulfur compounds. Calculate the enthalpy change of the following reaction.

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

Data: 　　　　$S(s) + O_2(g) \rightarrow SO_2(g)$ 　　　　　　　$\Delta H = -296.8 \text{ kJ mol}^{-1}$

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

Using $\Delta_r H° = \Sigma\, m\Delta_f H°$(products) - $\Sigma\, m\Delta_f H°$(reactants), the enthalpy change for the reaction can be written as

$\Delta_r H° = 2\Delta_f H°(SO_3(g)) - 2\Delta_f H°(SO_2(g))$

As $O_2(g)$ is an element in its standard state, its $\Delta_f H° = 0$.

The two reactions for which data are given correspond to the formation of $SO_2(g)$ and *two* moles $SO_3(g)$, respectively, from the elements in their standard states. Hence,

$\Delta_f H°(SO_2(g)) = -296.8 \text{ kJ mol}^{-1}$
$2\Delta_f H°(SO_3(g)) = -791.4 \text{ kJ mol}^{-1}$ or $\Delta_f H°(SO_3(g)) = -395.7$

Hence,

$\Delta_r H° = (2 \times -395.7) - (2 \times -296.8) \text{ kJ mol}^{-1} = -197.8 \text{ kJ mol}^{-1}$

Answer: **-197.8 kJ mol^{-1}**

1

- If 78.2 J is required to raise the temperature of 45.6 g of lead by 13.3 °C, what is the specific heat of lead in $J \text{ g}^{-1} \text{ K}^{-1}$?

The heat, q, required to change the temperature by an amount ΔT is given by:

$q = cm\Delta T$

where m is the mass and c is the specific heat capacity.

Hence as the temperature change is 13.3 °C or 13.3 K,

78.2 J = $c \times$ (45.6 g) $\times$ (13.3 K)　　or c = 0.129 J g^{-1} K^{-1}

Answer:

ANSWER CONTINUES ON THE NEXT PAGE

- Rank the following compounds in order of increasing boiling point? Justify your answer.

    $CH_3CH_2OCH_2CH_3$, CH_3OH, CH_4, CH_3CH_3, CH_3CH_2OH

3

CH_4 < CH_3CH_3 < $CH_3CH_2OCH_2CH_3$ < CH_3OH < CH_3CH_2OH

- **CH_4 and CH_3CH_3 have only weak dispersion forces; CH_3CH_3 has more electrons so has the dispersion forces are stronger.**

- **$CH_3CH_2OCH_2CH_3$ has dispersion forces and, as the presence of the electronegative oxygen atom leads to a dipole moment, dipole-dipole forces.**

- **CH_3CH_2OH and CH_3OH have H attached to electronegative O atoms and so have relatively strong H-bonds as well as dispersion forces and dipole-dipole forces. CH_3CH_2OH has more electrons so has stronger dispersion forces and hence the higher boiling point.**

**Marks
6**

- When a 1.00 g sample of carbon is burnt in a calorimeter to produce $CO_2(g)$, a temperature rise of 6.66 ºC is observed. When a 1.00 g sample of solid NH_4NO_3 is decomposed in the same calorimeter, a temperature rise of 0.300 ºC is observed. The equation for this reaction is:
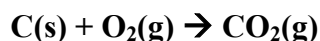
$$2NH_4NO_3(s) \rightarrow 2N_2(g) + 4H_2O(g) + O_2(g)$$
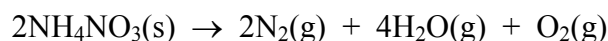
What is the heat of reaction for the decomposition of 1.00 kg of ammonium nitrate?

Heat of formation data: $\Delta_fH = -393.3$ kJ mol^{-1} for $CO_2(g)$

The combustion of carbon follow s the reaction,

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

Using $\Delta_rH° = \Sigma\ m\Delta_fH°$(products) - $\Sigma\ m\Delta_fH°$(reactants), the enthalpy change for the reaction can be written as

  $\Delta_rH° = \Delta_fH°(CO_2(g)) = -393.3$ kJ mol^{-1}

As $O_2(g)$ and $C(s)$ are an element in its standard state, $\Delta_fH°$(reactants) = 0.

1.00 g of carbon corresponds to $\dfrac{\text{mass}}{\text{atomic mass}} = \dfrac{1.00\ \text{g}}{12.01\ \text{g mol}^{-1}} = 0.0833$ mol.

As the combustion of 1 mole of $C(s)$ produces -393.3 kJ, the combustion of 0.0833 mol produces (0.0833 mol) × (-393.3 kJ mol^{-1}) = -32.7 kJ.

This produces a temperature _change_ in the calorimeter of 6.66 °C = 6.66 K. As $q = C\,\Delta T$, the heat capacity of the calorimeter is:

  C(calorimeter) = $q\ /\ \Delta T$ = (32.7 × 10^3 J) / (6.66 K) = 4920 J K^{-1}.

The combustion of 1.00 g of NH_4NO_3 produces a temperature change of 0.300 °C = 0.300 K. The corresponding heat change is therefore,

  $q = C\Delta T$ = (4902 J K^{-1}) × (0.300 K) = 1480 J

As a temperature _rise_ is found, the reaction must be exothermic and so, for the 1.00 g sample, Δ_rH = -1480 J.

Therefore, for 1.00 kg, the heat of reaction is,

  Δ_rH = -1480 J × 1000 = -1480 kJ

Answer: **-1480 kJ**

ANSWER CONTINUES ON THE NEXT PAGE

List all of the nitrogen containing species in this reaction. Beside each, give the oxidation number of the nitrogen in that species.

Nitrogen is contained in NH_4NO_3 and N_2.

N_2 contains nitrogen with an oxidation number of zero: 0.

NH_4NO_3 is made up of NH_4^+ and NO_3^-.

- **NH_4^+ contains $4H^+$ and so nitrogen has an oxidation number of –III.**
- **NO_3^- contains $3O^{2-}$ and so nitrogen has an oxidation number of +V**