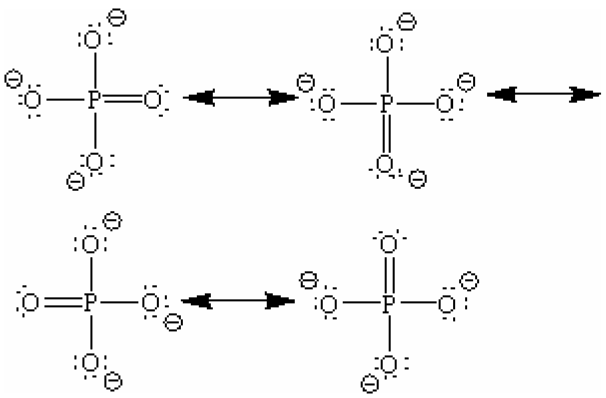
Marks

- Balance the following nuclear reactions by identifying the missing nuclear particle. **2**

$^{14}_{6}C$	$\rightarrow$	$\mathbf{^{14}_{7}N}$	+	$^{0}_{-1}e$
$^{243}_{95}Am$	$\rightarrow$	$\mathbf{^{239}_{93}Np}$	+	$^{4}_{2}He$

- Briefly explain the concept of resonance in Lewis structures. Include an example of a species that displays resonance. **3**

When more than one Lewis structure can be drawn for a molecule, the true structure is an average (called a resonance hybrid) of all the structures drawn. In some cases one or more structures may be dominant contributors - in other cases there may be only 2 equal contributors. For example, in the PO_4^{3} ion, there are four equivalent contributors:



- What element has the ground state electronic arrangement of $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^1$? **1**

Aluminium

- Calculate the frequency and energy of yellow light of wavelength 570 nm. **2**

A wavelength, λ, of 570 nm corresponds to 570 × 10^{-9} m.

The energy is given by $E = \dfrac{hc}{\lambda} = \dfrac{(6.634\times10^{-34})\times(2.998\times10^{8})}{(570\times10^{-9})} = 3.49\times10^{-19}$ **J**

The frequency, ν, is related to the wavelength by c = λν or

$\nu = \dfrac{c}{\lambda} = \dfrac{(2.998\times10^{8})}{(570\times10^{-9})} = 5.26\times10^{14}$ **Hz**

Frequency = **5.26 × 10^{14} Hz**	Energy = **3.49 × 10^{-19} J**

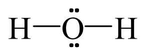
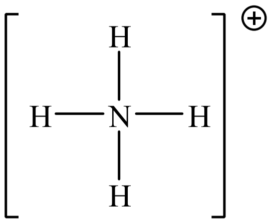
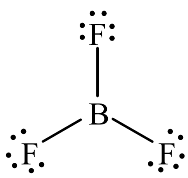
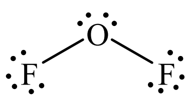
Marks

- Iodine is a soft, low-melting point solid while diamond is very hard and has an extremely high melting point. How does the bonding in each element account for these differences in properties? **3**

Iodine consists of discrete I_2 molecules. The intermolecular forces between these I_2 units are weak dispersion forces, so the solid is soft with a low melting point. The strength of the covalent I-I bond is essentially irrelevant.

Diamond consists of a giant 3-dimensional array of carbon atoms in a tetrahedral arrangement. Each atom is covalently bonded to its neighbour to give one giant molecule (covalent network solid). The C-C covalent bond is very strong, so diamond is hard with a high melting point.

- Draw Lewis diagrams of the following species. Give the arrangement of the valence shell electron pairs (both bonding and, where present, non-bonding) around the underlined atom and predict the geometry of each species. **6**

Species	Lewis diagram	Arrangement of electron pairs	Geometry of species
H_2<u>O</u>		tetrahedral	bent
<u>N</u>H_4^+		**tetrahedral**	**tetrahedral**
<u>B</u>F_3		**trigonal planar**	**trigonal planar**
<u>O</u>F_2		**tetrahedral**	**bent**

Marks

- The element boron forms a series of hydrides, which includes B_2H_6, B_4H_{10}, B_5H_9, B_6H_{10} and $B_{10}H_{14}$. Which one of these hydrides consists of 85.63% boron by mass? **2**

The molar mass of the boranes are:

molar mass of B_2H_6 = (2 × 10.81 (B)) + (6 × 1.008 (H)) = 27.668 g mol^{-1}
molar mass of B_4H_{10} = (4 × 10.81 (B)) + (10 × 1.008 (H)) = 53.32 g mol^{-1}
molar mass of B_5H_9 = (5 × 10.81 (B)) + (9 × 1.008 (H)) = 63.122 g mol^{-1}
molar mass of B_6H_{10} = (6 × 10.81 (B)) + (10 × 1.008 (H)) = 74.94 g mol^{-1}
molar mass of $B_{10}H_{14}$ = (10 × 10.81 (B)) + (14 × 1.008 (H)) = 122.212 g mol^{-1}

$$\textbf{The percentage of boron} = \frac{\text{mass of boron in one mole of hydride (in g)}}{\text{molar mass of boron hydride (in g mol}^{-1}\text{)}} \times 100\%$$

$$\text{percentage boron in } B_2H_6 = \frac{2 \times 10.81}{27.668} \times 100\% = 78.14\%$$

$$\text{percentage boron in } B_4H_{10} = \frac{4 \times 10.81}{53.32} \times 100\% = 81.10\%$$

$$\text{percentage boron in } B_5H_9 = \frac{5 \times 10.81}{63.122} \times 100\% = 85.63\%$$

$$\text{percentage boron in } B_6H_{10} = \frac{6 \times 10.81}{74.94} \times 100\% = 86.55\%$$

$$\text{percentage boron in } B_{10}H_{14} = \frac{10 \times 10.81}{122.12} \times 100\% = 88.45\%$$

Answer: **B_5H_9**

- Complete the following table. **2**

Formula	Name
K_2SO_4	**potassium sulfate**
$CuCl_2$	copper(II) chloride
SF_4	**sulfur(IV) fluoride**
K_2CrO_4	potassium chromate

Marks

- Balance the following equation: **2**

$$Fe_2O_3(s) + CO(g) \rightarrow Fe(s) + CO_2(g)$$

$Fe_2O_3(s) + 3CO(g) \rightarrow 2Fe(s) + 3CO_2(g)$

- Calculate the mass of sodium hydroxide required to make 500 mL of a 0.200 M aqueous solution. **6**

The number of moles in a solution is given by

number of moles = concentration (in M) × volume (in L)

The solution contains 0.2 × 500/1000 = 0.1 moles of NaOH.

The molar mass of NaOH is 22.99 (Na) + 16.00 (O) + 1.01 = 40.00 g mol^{-1}.

The mass of NaOH required is 0.1 × 40.00 = 4.00 g

Answer: **4.00 g**

What volume of the above solution would be required to neutralise 50.0 mL of 0.100 M hydrochloric acid solution?

The neutralization is a 1:1 reaction:

$HCl + NaOH \rightarrow NaCl + H_2O$

50.0 mL of 0.1000 HCl contains 0.1 × 50/1000 = 0.005 moles

The volume is given by

volume (in L) = number of moles (in mol) / concentration

The volume of the 0.200 M NaOH containing 0.005 moles is therefore:

volume = 0.005 / 0.200 = 0.025 L or 25 mL

Answer: **25.0 mL**

Marks

- A 0.50 g sample of ammonium nitrate, $NH_4NO_3(s)$, was dissolved in 35.0 g of water in a coffee cup calorimeter. The temperature of the solution dropped from 22.7 to 21.6 °C. Write a balanced equation to describe the reaction in the calorimeter. **5**

$$NH_4NO_3(s) \rightarrow NH_4^+(aq) + NO_3^-(aq)$$

Describe this process as either endothermic or exothermic.	**Temperature decreases so endothermic**

Assuming a perfect calorimeter what is the heat of solution of ammonium nitrate, expressed in kJ mol^{-1}? Assume the density of the solution is 1.00 g mL^{-1} and that the heat capacity of the solution is 4.18 J K^{-1} g^{-1}.

The molar mass of NH_4NO_3 is 14.01 (N) + 4 × 1.008 (H) + 14.01 (N) + 3 × 16.00 (O) = 80.052. The sample of 0.50 g therefore corresponds to:

$$\text{number of moles} = \frac{0.50}{80.052} = 0.0062 \text{ mol}$$

The total mass of NH_4NO_3 and water is 0.50 + 35.0 = 35.5 g. The heat change is given by:

$$q = c \times m \times \Delta T = 4.18 \times 35.5 \times (22.7 - 21.6) = 163 \text{ J or } 0.163 \text{ kJ}$$

This is the heat change produced by 0.0062 mol. The heat change produced by 1 mol is therefore $\frac{0.163}{0.0062}$ **= 26.1 kJ**

The reaction is endothermic as the temperature drops. The enthalpy change is therefore:

$$\Delta H_r = +26.1 \text{ kJ mol}^{-1}$$

- Heat radiating fins are used to dissipate heat and prevent damage to electronic components. Is it better to make the fins out of aluminium or iron? Give reasons for your answer. **2**
Data: Specific heat of Al = 0.900 J K^{-1} g^{-1} Specific heat of Fe = 0.444 J K^{-1} g^{-1}

The fins are required to remove heat from the electronic components. Aluminium has a higher heat capacity so it can absorb more heat per gram.

Marks

- Aluminium acts as a reducing agent in the thermite reaction where Fe_2O_3 is reduced to metallic iron. Write a balanced equation for the thermite reaction. **4**

$2Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2Fe(s)$

What is the maximum theoretical mass of Fe that can be produced when 270 g of Al reacts with excess Fe_2O_3 in the thermite reaction?

270 g of aluminium corresponds to:

$$\textbf{number of moles} = \frac{\textbf{mass of Al}}{\textbf{atomic mass of Al}} = \frac{\textbf{270}}{\textbf{26.98}} = \textbf{10. mol}$$

From the chemical equation, 2 mol of Fe is produced for every 2 mol of Al consumed. Hence, 10.00 mol of Fe is the maximum yield. This corresponds to:

mass of Fe = number of moles of Fe × molar mass of Fe
= 10. g × 55.86 g mol^{-1} = 560 g

Answer: **560 g**

- What does the superscript "o" mean in the symbol $\Delta H_f°$? **1**

All reactants and products are in their standard states: the most stable form of the substance at a pressure of 1 atm and a temperature of 298 K.

- Briefly describe what is meant by "Dynamic Equilibrium"? **1**

A reaction at equilibrium has not stopped - the rate of the forward reaction is equal to the rate of the backward reaction - a dynamic situation.

Marks 3

- A sealed 1.000 L flask at 30 °C contains air at a pressure of 1.000 atm. A 5.00 g sample of liquid water is injected into the flask and the flask heated to a temperature of 150 °C, causing the water to vaporise. What is the final pressure in the flask?

The pressure inside the flask will increase due to *both* the air and the vaporised water.

H_2O has a molar mass of 2×1.008 (H) + 16.00 (O) = 180.016. The number of moles of water is therefore:

$$\text{moles of water} = \frac{\text{mass of water}}{\text{molar mass of water}} = \frac{5.00}{18.016} = 0.278\ \text{mol}$$

For a perfect gas, PV = nRT. The pressure due to the water is therefore:

partial pressure of water = *nRT* /*V*
$$= (0.278\ \text{mol})(0.08206\ \text{atm L mol}^{-1}\ \text{K}^{-1})((150 + 273)\ \text{K}) / (1.000\ \text{L})$$
$$= 9.64\ \text{atm}$$

Note the use of R = 0.08206 L atm K^{-1} mol^{-1} allows the use of V = 1.000 L and gives the answers in atmospheres.

The pressure due to the air already in the flask increases due to the increase in temperature. As the number of moles of air and the volume stays the same:

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ or } P_2 = \frac{P_1 T_2}{T_1} = \frac{(1.000) \times (150 + 273)}{(30 + 273)} = 1.40\ \text{atm}$$

(Alternatively, PV = nRT can be used to work out the number of moles of air present and then used again to find the pressure exerted at the higher temperature).

The total pressure is therefore:

P = (9.64 + 1.40) atm = 11.0 atm

Answer: **11.0 atm**

Marks 4

- Consider the following reaction.

$$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(l) \qquad \Delta E = -10909 \text{ kJ mol}^{-1}$$

A mixture of C_8H_{18} (10.00 g) and O_2 (30.00 g) is allowed to react. Assuming that the reaction goes to completion, how much energy will be produced?

The molar mass of C_8H_{18} is 8×12.01 (C) + 18×1.008 (H) = 114.224. The number of moles of C_8H_{18} = $\frac{10.00}{114.224} = 0.08755$ mol

The molar mass of O_2 is 2×16.00 = 32.00.

The number of moles of O_2 is therefore $\frac{30.00}{16.00} = 0.9375$.

25 moles of O_2 is required for every 2 moles of C_8H_{18} – 12.5 mol of O_2 is required for every 1 mol of C_8H_{18}. The ratio of O_2 : C_8H_{18} is actually $\frac{0.9375}{0.08755} = 10.71$ so O_2 is the limiting reagent.

10909 kJ is produced for every 25 mol of O_2 that reacts. As 0.9375 mol of O_2 are present, the energy produced is:

$$\Delta E = \frac{0.9375}{25} \times 10909 = 409.1\,\text{kJ}$$

Answer: **409.1 kJ**

3

- The half reactions describing the discharge of a silver-zinc cell are:

$$Zn(s) + 2OH^-(aq) \rightarrow ZnO(s) + H_2O(l) + 2e^-$$

$$Ag_2O(s) + H_2O(l) + 2e^- \rightarrow 2Ag(s) + 2OH^-(aq)$$

List the chemical species that will be consumed as the battery discharges.

Zn(s) and Ag_2O(s)

Why is a saturated solution of KOH used in the battery?

It acts like a salt bridge allowing the migration of ions away from and towards the electrodes. OH^-(aq) ions are produced at the cathode and consumed at the anode.

Why is the voltage in the silver-zinc cell constant during discharge?

The overall cell reaction is:

$Zn(s) + Ag_2O(s) \rightarrow 2Ag(s) + ZnO(s)$.

No ions with constantly changing concentrations appear, so the cell produces a constant voltage until one of the reactants is exhausted and it stops functioning.

Marks

- Consider the following cell reaction. **3**

$$Cr^{2+}(aq) + Zn(s) \rightleftharpoons Cr(s) + Zn^{2+}(aq)$$

Use the Nernst equation to calculate the ratio of cation concentrations at 298 K for which the cell potential, E = 0 V.

The half reactions and potentials are:

$Cr^{2+}(aq) + 2e^- \rightarrow Cr(s)$	$E^\circ = -0.89$ V
$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$	$E^\circ = +0.76$ V (reversed for oxidation required)

Hence, $E^\circ = (-0.89) + (+0.76) = -0.13$ V

The Nernst equation gives $E_{cell} = E^0 - 2.303 \times \frac{RT}{nF}\log(Q)$. $E_{cell} = 0$ **when:**

$$E^0 = 2.303 \times \frac{RT}{nF}\log(Q) = 2.303 \times \frac{RT}{nF}\log(K_{eq})$$

The ratio of cation concentrations is just $Q = \frac{[Zn^{2+}(aq)]}{[Cr^{2+}(aq)]}$. **The process involves** $2e^-$ **so n = 2 and Q can be obtained using the** $E^\circ = -0.13$ **V:**

$$Q = 10^{nFE^\circ/2.303RT} = 10^{(2\times96485\times-0.13)/(2.303\times8.314\times298)} = 4\times10^{-5}$$

Answer: $\mathbf{4 \times 10^{-5}}$

- A lead-acid battery has the following shorthand notation: **4**

$Pb(s), PbSO_4(s) \mid H^+(aq), SO_4^{2-}(aq) \parallel H^+(aq), SO_4^{2-}(aq) \mid PbO_2(s), Pb(s)$

Which component of the battery is the anode? **$Pb(s), PbSO_4(s)$**

Give the balanced half equation of the reaction that takes place at the anode.

$Pb(s) + SO4^{2-}(aq) \rightarrow PbSO4(s) + 2e^-$

Which component of the battery is the cathode? **$PbO_2(s), PbSO_4(s)$**

Give the balanced half equation of the reaction that takes place at the cathode.

$PbO_2(s) + 4H^+(aq) + SO_4^{2-}(aq) + 2e^- \rightarrow PbSO_4(s) + 2H_2O$

Marks
4

- At 800 °C, the value of the equilibrium constant, K_c, for the following equation is 1.245×10^3 L mol^{-1}.

$$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$$

What is the equilibrium concentration of NO(g) at 800 °C if, at equilibrium, $[O_2(g)] = 0.0012$ M and $[NO_2(g)] = 0.055$ M?

The equilibrium constant in terms of concentrations, K_c, is given by:

$$K_c = \frac{[NO_2(g)]^2}{[NO(g)]^2[O_2(g)]}$$

Using $[O_2(g)] = 0.0012$ M, $[NO_2(g)] = 0.055$ M and $K_c = 1.245 \times 10^3$:

$$K_c = \frac{(0.055)^2}{[NO(g)]^2 \times (0.0012)} = 1.245 \times 10^3$$

so

$$[NO(g)]^2 = \frac{(0.055)^2}{(1.245 \times 10^3) \times (0.0012)} = 0.0020\,M^2$$

$$[NO(g)]^2 = \sqrt{0.0020} = 0.045\,M$$

Answer: **0.045 M**