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

\[
\begin{align*}
^{14}_{6}\text{C} & \rightarrow ^{14}_{7}\text{N} & + & ^{0}_{-1}\text{e} \\
^{243}_{95}\text{Am} & \rightarrow ^{239}_{93}\text{Np} & + & ^{4}_{2}\text{He}
\end{align*}
\]

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

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 \(\text{PO}_4^{3-}\) ion, there are four equivalent contributors:

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

Aluminium

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

A wavelength, \(\lambda\), of 570 nm corresponds to \(570 \times 10^{-9} \text{ m}\).

The energy is given by 
\[
E = \frac{hc}{\lambda} = \frac{(6.634 \times 10^{-34}) \times (2.998 \times 10^8)}{(570 \times 10^{-9})} = 3.49 \times 10^{-19} \text{ J}
\]

The frequency, \(\nu\), is related to the wavelength by 
\[
\nu = \frac{c}{\lambda} = \frac{(2.998 \times 10^8)}{(570 \times 10^{-9})} = 5.26 \times 10^{14} \text{ Hz}
\]

Frequency = \(5.26 \times 10^{14} \text{ Hz}\)  |  Energy = \(3.49 \times 10^{-19} \text{ J}\)
• 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?

Iodine consists of discrete I₂ molecules. The intermolecular forces between these I₂ 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.

<table>
<thead>
<tr>
<th>Species</th>
<th>Lewis diagram</th>
<th>Arrangement of electron pairs</th>
<th>Geometry of species</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₂O</td>
<td>( \text{H} \overset{\cdot}{\text{O}} \overset{\cdot}{\text{H}} )</td>
<td>tetrahedral</td>
<td>bent</td>
</tr>
</tbody>
</table>
| NH₄⁺    | \[
\begin{array}{c}
\text{H} \\
\text{H} \\
\text{H} \\
\text{H} \\
\text{N} \\
\text{H}
\end{array}
\] | tetrahedral                  | tetrahedral          |
| BF₃     | \[
\begin{array}{c}
\text{F} \\
\text{B} \\
\text{F} \\
\text{F} \\
\text{F}
\end{array}
\] | trigonal planar              | trigonal planar       |
| OF₂     | \[
\begin{array}{c}
\text{F} \\
\text{O} \\
\text{F} \\
\text{F}
\end{array}
\] | tetrahedral                  | bent                 |
The element boron forms a series of hydrides, which includes $\text{B}_2\text{H}_6$, $\text{B}_4\text{H}_{10}$, $\text{B}_5\text{H}_9$, $\text{B}_6\text{H}_{10}$ and $\text{B}_{10}\text{H}_{14}$. Which one of these hydrides consists of 85.63% boron by mass?

The molar mass of the boranes are:

- Molar mass of $\text{B}_2\text{H}_6 = (2 \times 10.81 \text{ (B)}) + (6 \times 1.008 \text{ (H)}) = 27.668 \text{ g mol}^{-1}$
- Molar mass of $\text{B}_4\text{H}_{10} = (4 \times 10.81 \text{ (B)}) + (10 \times 1.008 \text{ (H)}) = 53.32 \text{ g mol}^{-1}$
- Molar mass of $\text{B}_5\text{H}_9 = (5 \times 10.81 \text{ (B)}) + (9 \times 1.008 \text{ (H)}) = 63.122 \text{ g mol}^{-1}$
- Molar mass of $\text{B}_6\text{H}_{10} = (6 \times 10.81 \text{ (B)}) + (10 \times 1.008 \text{ (H)}) = 74.94 \text{ g mol}^{-1}$
- Molar mass of $\text{B}_{10}\text{H}_{14} = (10 \times 10.81 \text{ (B)}) + (14 \times 1.008 \text{ (H)}) = 122.212 \text{ g mol}^{-1}$

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})} \times 100\% \ \\

\begin{align*}
\text{percentage boron in } \text{B}_2\text{H}_6 & = \frac{2 \times 10.81}{27.668} \times 100\% = 78.14\% \\
\text{percentage boron in } \text{B}_4\text{H}_{10} & = \frac{4 \times 10.81}{53.32} \times 100\% = 81.10\% \\
\text{percentage boron in } \text{B}_5\text{H}_9 & = \frac{5 \times 10.81}{63.122} \times 100\% = 85.63\% \\
\text{percentage boron in } \text{B}_6\text{H}_{10} & = \frac{6 \times 10.81}{74.94} \times 100\% = 86.55\% \\
\text{percentage boron in } \text{B}_{10}\text{H}_{14} & = \frac{10 \times 10.81}{122.212} \times 100\% = 88.45\%
\end{align*}

Answer: $\text{B}_5\text{H}_9$

- Complete the following table.

<table>
<thead>
<tr>
<th>Formula</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{K}_2\text{SO}_4$</td>
<td>potassium sulfate</td>
</tr>
<tr>
<td>$\text{CuCl}_2$</td>
<td>copper(II) chloride</td>
</tr>
<tr>
<td>$\text{SF}_4$</td>
<td>sulfur(IV) fluoride</td>
</tr>
<tr>
<td>$\text{K}_2\text{CrO}_4$</td>
<td>potassium chromate</td>
</tr>
</tbody>
</table>
• Balance the following equation:

\[
\text{Fe}_2\text{O}_3(s) + \text{CO(g)} \rightarrow \text{Fe(s)} + \text{CO}_2(g)
\]

\[
\text{Fe}_2\text{O}_3(s) + 3\text{CO(g)} \rightarrow 2\text{Fe(s)} + 3\text{CO}_2(g)
\]

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

The number of moles in a solution is given by

\[
\text{number of moles} = \text{concentration (in M)} \times \text{volume (in L)}
\]

The solution contains 0.2 \times \frac{500}{1000} = 0.1 \text{ 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 \times 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:

\[
\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}
\]

50.0 mL of 0.1000 HCl contains 0.1 \times \frac{50}{1000} = 0.005 \text{ moles}

The volume is given by

\[
\text{volume (in L)} = \frac{\text{number of moles (in mol)}}{\text{concentration}}
\]

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

\[
\text{volume} = \frac{0.005}{0.200} = 0.025 \text{ L or } 25 \text{ mL}
\]

Answer: 25.0 mL
A 0.50 g sample of ammonium nitrate, $\text{NH}_4\text{NO}_3$ (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.

\[
\text{NH}_4\text{NO}_3 (s) \rightarrow \text{NH}_4^+ (aq) + \text{NO}_3^- (aq)
\]

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 $\text{NH}_4\text{NO}_3$ 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 $\text{NH}_4\text{NO}_3$ and water is 0.50 + 35.0 = 35.5 g. The heat change is given by:

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

This is the heat change produced by 0.0062 mol. The heat change produced by 1 mol is therefore

\[
\frac{0.163}{0.0062} = 26.1 \text{ kJ}
\]

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.

Data: Specific heat of $\text{Al} = 0.900 \text{ J K}^{-1} \text{ g}^{-1}$ Specific heat of $\text{Fe} = 0.444 \text{ J K}^{-1} \text{ g}^{-1}$

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.
Aluminium acts as a reducing agent in the thermite reaction where Fe₂O₃ is reduced to metallic iron. Write a balanced equation for the thermite reaction.

\[ 2\text{Al}(s) + \text{Fe}_2\text{O}_3(s) \rightarrow \text{Al}_2\text{O}_3(s) + 2\text{Fe}(s) \]

What is the maximum theoretical mass of Fe that can be produced when 270 g of Al reacts with excess Fe₂O₃ in the thermite reaction?

270 g of aluminium corresponds to:

\[
\text{number of moles} = \frac{\text{mass of Al}}{\text{atomic mass of Al}} = \frac{270}{26.98} = 10.00 \text{ mol}
\]

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:

\[
\text{mass of Fe} = \text{number of moles of Fe} \times \text{molar mass of Fe} = 10.00 \text{ g} \times 55.86 \text{ g mol}^{-1} = 560 \text{ g}
\]

Answer: 560 g

What does the superscript “°” mean in the symbol Δ\(H_f\)?

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”?

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.
• 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.

H₂O has a molar mass of 2×1.008 (H) + 16.00 (O) = 18.016. 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:

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

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:

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

(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:

\[
P = (9.64 + 1.40) \text{ atm} = 11.0 \text{ atm}
\]

Answer: 11.0 atm
Consider the following reaction.

$$2\text{C}_8\text{H}_{18}(l) + 25\text{O}_2(g) \rightarrow 16\text{CO}_2(g) + 18\text{H}_2\text{O}(l) \quad \Delta E = -10909 \text{ kJ mol}^{-1}$$

A mixture of $\text{C}_8\text{H}_{18}$ (10.00 g) and $\text{O}_2$ (30.00 g) is allowed to react. Assuming that the reaction goes to completion, how much energy will be produced?

The molar mass of $\text{C}_8\text{H}_{18}$ is $8 \times 12.01 \text{ (C)} + 18 \times 1.008 \text{ (H)} = 114.224$. The number of moles of $\text{C}_8\text{H}_{18} = \frac{10.00}{114.224} = 0.08755 \text{ mol}$

The molar mass of $\text{O}_2$ is $2 \times 16.00 = 32.00$.

The number of moles of $\text{O}_2$ is therefore $\frac{30.00}{16.00} = 0.9375$.

25 moles of $\text{O}_2$ is required for every 2 moles of $\text{C}_8\text{H}_{18}$. 12.5 mol of $\text{O}_2$ is required for every 1 mol of $\text{C}_8\text{H}_{18}$. The ratio of $\text{O}_2 : \text{C}_8\text{H}_{18}$ is actually $\frac{0.9375}{0.08755} = 10.71$ so $\text{O}_2$ is the limiting reagent.

10909 kJ is produced for every 25 mol of $\text{O}_2$ that reacts. As 0.9375 mol of $\text{O}_2$ are present, the energy produced is:

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

Answer: 409.1 kJ

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

$$\text{Zn(s)} + 2\text{OH}^-(aq) \rightarrow \text{ZnO(s)} + \text{H}_2\text{O}(l) + 2\text{e}^-$$
$$\text{Ag}_2\text{O(s)} + \text{H}_2\text{O}(l) + 2\text{e}^- \rightarrow 2\text{Ag(s)} + 2\text{OH}^- (aq)$$

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

$\text{Zn(s)}$ and $\text{Ag}_2\text{O(s)}$

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$^-$ 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:

\[ \text{Zn(s)} + \text{Ag}_2\text{O(s)} \rightarrow 2\text{Ag(s)} + \text{ZnO(s)} \].

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.
• Consider the following cell reaction.

\[
\text{Cr}^{2+}(aq) + \text{Zn}(s) \rightleftharpoons \text{Cr}(s) + \text{Zn}^{2+}(aq)
\]

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:

\[
\begin{align*}
\text{Cr}^{2+}(aq) + 2e^- &\rightarrow \text{Cr}(s) \quad E^\circ = -0.89 \text{ V} \\
\text{Zn}(s) &\rightarrow \text{Zn}^{2+}(aq) + 2e^- \quad E^\circ = +0.76 \text{ V (reversed for oxidation required)}
\end{align*}
\]

Hence, \( E^\circ = (-0.89) + (+0.76) = -0.13 \text{ V} \)

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

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

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

\[
Q = 10^{nFE^0/2.303RT} = 10^{(2 \times 96485 \times -0.13)/(2.303 \times 8.314 \times 298)} = 4 \times 10^{-5}
\]

Answer: \( 4 \times 10^{-5} \)

• A lead-acid battery has the following shorthand notation:

\[
\text{Pb}(s), \text{PbSO}_4(s) | \text{H}^+(aq), \text{SO}_4^{2-}(aq) || \text{H}^+(aq), \text{SO}_4^{2-}(aq) | \text{PbO}_2(s), \text{Pb}(s)
\]

Which component of the battery is the anode? \( \text{Pb}(s), \text{PbSO}_4(s) \)

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

\[
\text{Pb}(s) + \text{SO}_4^{2-}(aq) \rightarrow \text{PbSO}_4(s) + 2e^- 
\]

Which component of the battery is the cathode? \( \text{PbO}_2(s), \text{PbSO}_4(s) \)

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

\[
\text{PbO}_2(s) + 4\text{H}^+(aq) + \text{SO}_4^{2-}(aq) + 2e^- \rightarrow \text{PbSO}_4(s) + 2\text{H}_2\text{O}
\]
At 800 °C, the value of the equilibrium constant, $K_c$, for the following equation is $1.245 \times 10^3$ L mol$^{-1}$.

$$2\text{NO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{NO}_2(g)$$

What is the equilibrium concentration of NO(g) at 800 °C if, at equilibrium, $[\text{O}_2(g)] = 0.0012$ M and $[\text{NO}_2(g)] = 0.055$ M?

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

$$K_c = \frac{[\text{NO}_2(g)]^2}{[\text{NO}(g)]^2[\text{O}_2(g)]}$$

Using $[\text{O}_2(g)] = 0.0012$ M, $[\text{NO}_2(g)] = 0.055$ M and $K_c = 1.245 \times 10^3$:

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

so

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

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

Answer: 0.045 M