Calculations [S]

1. Ammonium nitrate (NH_4NO_3) is used widely as a fertiliser and is made from the reaction between ammonium hydroxide and nitric acid solutions:

   $NH_4OH(aq) + HNO_3(aq) \rightarrow NH_4NO_3(aq) + H_2O(l)$

   a. Calculate the relative formula mass (RFM) of nitric acid: **[1]**

   b. Calculate the amount in moles of 1.6g of ammonium nitrate: **[2]**

   c. Calculate the percentage by mass of nitrogen in ammonium nitrate: **[2]**

   d. For each batch of 4kg of ammonium hydroxide used:

      i. Calculate the mass in kg of ammonium nitrate produced: **[3]**

      ii. Calculate the mass of nitrogen contained within this ammonium nitrate: **[2]**

2. The formula of iron oxide can be calculated by heating iron in a crucible and measuring the mass of oxygen that combines with a fixed mass of iron. In this experiment, 3.92g of iron increased in mass to 5.60g when combined with oxygen.

    a. Define the term *empirical formula*: **[2]**

    b. Calculate the mass of oxygen that has reacted: **[1]**

    c. Determine the empirical formula of iron oxide: **[3]**

    d. In a separate experiment, the empirical formula of a hydrocarbon was found to be C_3H_7 and its RFM was 86. Determine its molecular formula: **[2]**

3. The reaction between sodium metal and water produces sodium hydroxide solution and hydrogen gas according to the following equation:

   $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

   a. The relative atomic mass of sodium is 23.0. Define the term *relative atomic mass*: **[2]**

   b. Calculate the mass of sodium needed to produce $75cm^3$ of hydrogen gas: **[3]**

   c. The sodium hydroxide produced is dissolved in the solution, whose volume is $250cm^3$.

      i. Calculate the mass of sodium hydroxide produced if 1.38g of sodium is used: **[3]**

      ii. Calculate the concentration of the sodium hydroxide solution that forms: **[2]**

   d. Calculate how many hydrogen ions are present in $1cm^3$ of 0.4 mol/dm^3 HCl(aq): **[2]**

4. Crystals of aluminium sulphate contain trapped molecules of water that can be removed by strong heating according to the following equation:

$Al_2SO_4{\cdot}xH_2O(s) \rightarrow Al_2SO_4(s) + xH_2O(l)$

   a. Ideally, after 0.510g of crystals is heated strongly 0.375g of powder remains.

      i. Calculate the amount in moles of water lost: **[2]**

      ii. Calculate the amount in moles of powder remaining: **[1]**

      iii. Deduce the value of *x*, the water of crystallisation of aluminium sulphate: **[2]**

   b. In fact, when 0.510g of crystals is heated strongly only 0.110g of water is lost. Some of the water remains trapped in the crystals.

      i. Calculate the mass of powder that would remain if this much water is lost: **[3]**

      ii. Calculate the percentage yield of powder: **[2]**

Calculations [S]

1. Ammonium nitrate (NH_4NO_3) is used widely as a fertiliser and is made from the reaction between ammonium hydroxide and nitric acid solutions:

   $NH_4OH(aq)$ + $HNO_3(aq)$ → $NH_4NO_3(aq)$ + $H_2O(l)$

   a. Calculate the relative formula mass (RFM) of nitric acid: **[1]**

      RFM = 1 + 14 + 3(16) = 63 [1]

   b. Calculate the amount in moles of 1.6g of ammonium nitrate: **[2]**

      RFM = 14 + 4(1) + 14 + 3(48) = 80 [1]

      moles = mass/RFM = 1.6/80 = 0.02mol [1]

   c. Calculate the percentage by mass of nitrogen in ammonium nitrate: **[2]**

      mass of N = 2(14) = 28 [1, or implied by correct next bit]

      % mass = 28/80 = 35% [1]

   d. For each batch of 4kg of ammonium hydroxide used:

      i. Calculate the mass in kg of ammonium nitrate produced: **[3]**

         moles NH_4OH = 4000/35 = 114.2857mol [1]

         moles NH_4NO_3 = 114.2857mol (1:1 ratio) [1]

         mass NH_4NO_3 = 114.2857 * 80 = 9142.857g = 9.14kg [1]

      ii. Calculate the mass of nitrogen contained within this ammonium nitrate: **[2]**

         mass N = 35/100 * 9142.857g [1]

         = 3200g or 3.2kg [1]

2. The formula of iron oxide can be calculated by heating iron in a crucible and measuring the mass of oxygen that combines with a fixed mass of iron. In this experiment, 3.92g of iron increased in mass to 5.60g when combined with oxygen.

   a. Define the term *empirical formula*: **[2]**

      The simplest… [1]

      … whole-number ratio of atoms/ions in a compound [1]

   b. Calculate the mass of oxygen that has reacted: **[1]**

      mass O 5.60 – 3.92 = 1.68g [1]

   c. Determine the empirical formula of iron oxide: **[3]**

Fe	**O**
3.92g	**1.68g**
3.92/56 = 0.07mol	**1.68/16 = 0.105mol**
Ratio is 2:3 so formula is Fe_2O_3	

[1] for each mole calculation

[1] for the final formula

   d. In a separate experiment, the empirical formula of a hydrocarbon was found to be C_3H_7 and its RFM was 86. Determine its molecular formula: **[2]**

      Number of units = RFM(proper)/RFM(empirical) = 2 [1]

      Molecular formula = 2 * C_3H_7 = C_6H_{14} [1]

3. The reaction between sodium metal and water produces sodium hydroxide solution and hydrogen gas according to the following equation:

$2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$

a. The relative atomic mass of sodium is 23.0. Define the term *relative atomic mass*: **[2]**

The average of the masses of its isotopes… [1]

… weighted by abundance. [1]

b. Calculate the mass of sodium needed to produce $75cm^3$ of hydrogen gas: **[3]**

moles H_2 = 75/24,000 = 0.003125mol [1]

moles Na = 2 * 0.003125 = 0.00625mol (2:1 ratio) [1]

mass Na = 0.00625 * 23 = 0.14375g (0.144g to 3sf) [1]

c. The sodium hydroxide produced is dissolved in the solution, whose volume is $250cm^3$.

i. Calculate the mass of sodium hydroxide produced if 1.38g of sodium is used: **[3]**

moles Na = 1.38/23 = 0.06mol [1]

moles NaOH = 0.06mol (1:1 ratio) [1]

mass NaOH = 0.06 * 40 = 2.4g [1]

ii. Calculate the concentration of the sodium hydroxide solution that forms: **[2]**

concentration = moles/volume = 0.06/(250/1000) [1]

= $0.24mol/dm^3$ [1]

d. Calculate how many hydrogen ions are present in $1cm^3$ of 0.4 mol/dm^3 HCl(aq): **[2]**

moles H^+ = concentration * volume = 0.4 * (1/1000) = 0.0004mol [1]

number = 0.0004 * $6x10^{23}$ = $2.4x10^{20}$ [1]

4. Crystals of aluminium sulphate contain trapped molecules of water that can be removed by strong heating according to the following equation:

   $Al_2SO_4 \cdot xH_2O(s) \rightarrow Al_2SO_4(s) + xH_2O(l)$

   a. Ideally, after 0.510g of crystals is heated strongly 0.375g of powder remains.

      i. Calculate the amount in moles of water lost: **[2]**

         mass H_2O = 0.510 – 0.375 = 0.135g [1]

         moles H_2O = 0.135/18 = 0.0075mol [1]

      ii. Calculate the amount in moles of powder remaining: **[1]**

         moles Al_2SO_4 = 0.375/150 = 0.0025mol [1]

      iii. Deduce the value of *x*, the water of crystallisation of aluminium sulphate: **[2]**

         x/1 = 0.0075/0.0025 (or equivalent ratio/fraction comparison) [1]

         x = 3 [1]

   b. In fact, when 0.510g of crystals is heated strongly only 0.110g of water is lost. Some of the water remains trapped in the crystals.

      i. Calculate the mass of powder that would remain if this much water is lost: **[3]**

         moles H_2O = 0.110/18 = 0.006111mol [1]

         moles Al_2SO_4 = 1/3 * 0.006111mol = 0.002037mol (1:3 ratio) [1]

         mass Al_2SO_4 = 0.002037 * 150 = 0.3056g [1]

      ii. Calculate the percentage yield of powder: **[2]**

         % yield = (actual/theoretical) * 100 = (0.3056/0.375)*100 [1]

         = 81.5% [1]