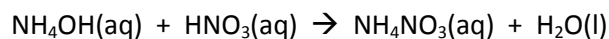


# Calculations [D]

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1. Ammonium nitrate ( $\text{NH}_4\text{NO}_3$ ) is used widely as a fertiliser and is made from the reaction between ammonium hydroxide and nitric acid solutions:



- 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.

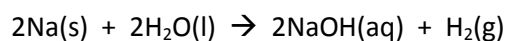
a. Define the term *empirical formula*: **[2]**

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:



- 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<sup>3</sup> of hydrogen gas: **[3]**
- c. The sodium hydroxide produced is dissolved in the solution, whose volume is 250cm<sup>3</sup>.
- 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]**

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

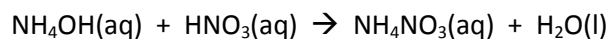


- a. Ideally, after 0.510g of crystals is heated strongly 0.375g of powder remains.
- Calculate the amount in moles of water lost: **[2]**
  - Calculate the amount in moles of powder remaining: **[1]**
  - Deduce the value of  $x$ , the water of crystallisation of aluminium sulphate: **[2]**
- b. In 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.
- Calculate the mass of powder that would remain if this much water is lost: **[3]**

# Calculations [D]

---

1. Ammonium nitrate ( $\text{NH}_4\text{NO}_3$ ) is used widely as a fertiliser and is made from the reaction between ammonium hydroxide and nitric acid solutions:



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

$$\text{RFM} = 1 + 14 + 3(16) = 63 \text{ [1]}$$

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

$$\text{RFM} = 14 + 4(1) + 14 + 3(48) = 80 \text{ [1]}$$

$$\text{moles} = \text{mass}/\text{RFM} = 1.6/80 = 0.02\text{mol} \text{ [1]}$$

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

$$\text{mass of N} = 2(14) = 28 \text{ [1, or implied by correct next bit]}$$

$$\% \text{ mass} = 28/80 = 35\% \text{ [1]}$$

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

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

$$\text{moles NH}_4\text{OH} = 4000/35 = 114.2857\text{mol} \text{ [1]}$$

$$\text{moles NH}_4\text{NO}_3 = 114.2857\text{mol} \text{ (1:1 ratio) [1]}$$

$$\text{mass NH}_4\text{NO}_3 = 114.2857 * 80 = 9142.857\text{g} = 9.14\text{kg} \text{ [1]}$$

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

$$\text{mass N} = 35/100 * 9142.857\text{g} \text{ [1]}$$

$$= 3200\text{g or } 3.2\text{kg} \text{ [1]}$$

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  $5.60 - 3.92 = 1.68\text{g}$  [1]**

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

<b>Fe</b>	<b>O</b>
<b>3.92g</b>	<b>1.68g</b>
<b><math>3.92/56 = 0.07\text{mol}</math></b>	<b><math>1.68/16 = 0.105\text{mol}</math></b>
<b>Ratio is 2:3 so formula is <math>\text{Fe}_2\text{O}_3</math></b>	

**[1] for each mole calculation**

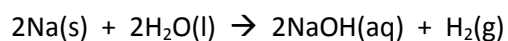
**[1] for the final formula**

d. In a separate experiment, the empirical formula of a hydrocarbon was found to be  $\text{C}_3\text{H}_7$  and its RFM was 86. Determine its molecular formula: [2]

**Number of units =  $\text{RFM}(\text{proper})/\text{RFM}(\text{empirical}) = 2$  [1]**

**Molecular formula =  $2 * \text{C}_3\text{H}_7 = \text{C}_6\text{H}_{14}$  [1]**

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



- 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<sup>3</sup> of hydrogen gas: [3]

$$\text{moles H}_2 = 75/24,000 = 0.003125\text{mol [1]}$$

$$\text{moles Na} = 2 * 0.003125 = 0.00625\text{mol (2:1 ratio) [1]}$$

$$\text{mass Na} = 0.00625 * 23 = 0.14375\text{g (0.144g to 3sf) [1]}$$

- c. The sodium hydroxide produced is dissolved in the solution, whose volume is 250cm<sup>3</sup>.

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

$$\text{moles Na} = 1.38/23 = 0.06\text{mol [1]}$$

$$\text{moles NaOH} = 0.06\text{mol (1:1 ratio) [1]}$$

$$\text{mass NaOH} = 0.06 * 40 = 2.4\text{g [1]}$$

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

$$\text{concentration} = \text{moles/volume} = 0.06/(250/1000) [1]$$

$$= 0.24\text{mol/dm}^3 [1]$$

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



- 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]**

$$\text{mass H}_2\text{O} = 0.510 - 0.375 = 0.135\text{g [1]}$$

$$\text{moles H}_2\text{O} = 0.135/18 = 0.0075\text{mol [1]}$$

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

$$\text{moles Al}_2\text{SO}_4 = 0.375/150 = 0.0025\text{mol [1]}$$

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

$$x/1 = 0.0075/0.0025 \text{ (or equivalent ratio/fraction comparison) [1]}$$

$$x = 3 \text{ [1]}$$

- 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]**

$$\text{moles H}_2\text{O} = 0.110/18 = 0.006111\text{mol [1]}$$

$$\text{moles Al}_2\text{SO}_4 = 1/3 * 0.006111\text{mol} = 0.002037\text{mol (1:3 ratio) [1]}$$

$$\text{mass Al}_2\text{SO}_4 = 0.002037 * 150 = 0.3056\text{g [1]}$$