- What is the molarity of the solution formed when 0.50 g of aluminium fluoride is dissolved in 800.0 mL of water?

Marks
2

The molar mass of AlF_3 is:

molar mass = (26.98 (Al) + 3 × 19.00 (F)) g mol^{-1} = 83.98 g mol^{-1}

The number of moles in 0.50 g is therefore:

number of moles = mass / molar mass = 0.50 g / 83.98 g mol^{-1} = 0.0060 mol

The concentration of this amount in 800.0 mL is then:

concentration = number of moles / volume
= 0.0060 mol / 0.8000 L = 0.0074 mol L^{-1}

Answer: **0.0074 mol L^{-1} or 0.0074 M**

What is [F^-] in this solution?

As the formula is AlF_3, dissolution results in 3F^-(aq) per formula unit.

[F^-(aq)] = 3 × 0.0074 mol L^{-1} = 0.022 mol L^{-1}

Answer: **0.022 mol L^{-1} or 0.022 M**

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

Marks
3

- In an experiment, 5.0 g of magnesium was dissolved in excess hydrochloric acid to give magnesium ions and hydrogen gas. Write a balanced equation for the reaction that occurred.

$Mg(s) + 2H^{+}(aq) \rightarrow Mg^{2+}(aq) + H_2(g)$

What amount of hydrogen gas (in mol) is produced in the reaction?

The molar mass of Mg is 24.31 g mol^{-1}. 5.0 g therefore corresponds to:

number of moles = mass / molar mass = 5.0 g / 24.31 g mol^{-1} = 0.21 mol

From the chemical equation, each mol of Mg that reacts will give one mol of H_2. Hence,

number of moles of H_2 = 0.21 mol.

Answer: **0.21 mol**

Marks 7

- A 0.060 M solution of aluminium nitrate and a 0.080 M solution of potassium phosphate are prepared by dissolving $Al(NO_3)_3$ and K_3PO_4 in water. Write the ionic equations for these two dissolutions reactions.

Dissolution of $Al(NO_3)_3$	$\mathbf{Al(NO_3)_3(s) \rightarrow Al^{3+}(aq) + 3NO_3^{-}(aq)}$
Dissolution of K_3PO_4	$\mathbf{K_3PO_4(s) \rightarrow 3K^{+}(aq) + PO_4^{3-}(aq)}$

If these solutions are combined, aluminium phosphate precipitates. Write the ionic equation for the precipitation reaction.

$\mathbf{Al^{3+}(aq) + PO_4^{3-}(aq) \rightarrow AlPO_4(s)}$

100.0 mL of the aluminium nitrate solution is added to 50.0 mL of the potassium phosphate solution. What amount (in mol) of aluminium phosphate precipitates?

100.0 mL of a 0.060 M solution of $Al(NO_3)_3$ contains:

number of moles of Al^{3+} = concentration × volume = $c \times V$
= 0.060 mol L^{-1} × 0.1000 L = 0.0060 mol

50.0 mL of a 0.080 M solution of K_3PO_4 contains:

number of moles of PO_4^{3-} = concentration × volume = $c \times V$
= 0.080 mol L^{-1} × 0.0500 L = 0.0040 mol

As the ionic equation has a 1 : 1 ratio of Al^{3+} : PO_4^{3-} reacting, PO_4^{3-} is the limiting reagent. The ionic equation shows that 1 mol of $AlPO_4$ is made from 1 mol of PO_4^{3-} so 0.0040 mol will produce 0.0040 mol.

Answer: **0.0040 mol**

What is the final concentration of aluminium ions remaining in solution after the precipitation?

Formation of 0.0040 mol of $AlPO_4$ requires 0.0040 mol of Al^{3+}. Therefore, the amount remaining is:

number of moles of Al^{3+} remaining = (0.0060 – 0.0040) mol = 0.0020 mol

After mixing the total solution volume is (100.0 + 50.0) mL = 150.0 mL. Hence, the concentration of Al^{3+}(aq) is:

concentration = number of moles / volume = n / V
= 0.0020 mol / 0.1500 L = 0.013 mol L^{-1}

Answer: **0.013 M**

- Explain why relative atomic masses are not always close to an integer. For example, copper has a reported value of 63.54. **2**

Many elements consist of isotopes, *i.e.* atoms with different numbers of neutrons and hence different atomic masses. The atomic mass of each isotope is close to an integer value. The relative atomic mass of an element is calculated using all these different isotopic masses and their relative percentages.

- Analysis of a black-coloured mineral called pitchblende returned the following percentage composition by weight: 84.80% uranium and 15.20% oxygen.
  What is the empirical formula of this compound? **2**

The mineral contains 84.80% U and 15.20% O.

	U	O
percentage	**84.80**	**15.20**
divide by atomic mass	$\frac{84.80}{238.03} = 0.356$	$\frac{15.20}{16.00} = 0.950$
divide by smallest value	**1**	**2.67**

The ratio of U : O is 1 : 2.67. The simplest whole number ratio can be obtained by multiplying this by 3 to give U : O equal to 3 : 8.

The empirical formula is U_3O_8.

Answer: $\mathbf{U_3O_8}$

Marks
3

- Balance the following equation:

$$NH_3(g) + O_2(g) \rightarrow NO(g) + H_2O(l)$$

$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(l)$

Calculate the mass of NH_3 required to produce 140. g of water.

The molar mass of H_2O is:

molar mass = [2 × 1.008 (H) + 16.00 (O)] g mol^{-1} = 18.016 g mol^{-1}

Hence, the number of moles of water produced is:

number of moles = mass / molar mass = (140. g) / (18.016 g mol^{-1}) = 7.771 mol

From the balanced equation, 4 mol of NH_3 will produce 6 mol of H_2O. Hence, to produce 7.771 mol of H_2O so:

number of moles of NH_3 = (4/6) × 7.771 mol = 5.18 mol

The molar mass of NH_3 is:

molar mass = [14.01 (N) + 3 × 1.008 (H)] g mol^{-1} = 17.034 g mol^{-1}

The mass of NH_3 in 5.18 mol is therefore:

mass = number of moles × molar mass = (5.18 mol) × (17.034 g mol^{-1}) = 88.2 g

Answer: **88.2 g**

- Calculate the number of aluminium atoms in a block of pure aluminium that measures 2.0 cm × 2.0 cm × 3.0 cm. The density of aluminium is 2.7 g cm^{-3}.

2

The volume of the block is:

V **= length × width × height = (2.0 × 2.0 × 3.0) cm^3 = 12 cm^3**

The mass can then be calculated from the density:

density = mass / volume or mass = density × volume

mass = (2.7 g cm^{-3}) × (12.0 cm^3) = 32.4 g

1 mol of Al has a mass equal to its atomic mass, 26.98 g mol^{-1} and contains 6.022 × 10^{23} mol^{-1}. Hence, the number of atoms in 32.4 g is:

number of atoms = number of moles × Avogadro's number
= (32.4 g / 26.98 g mol^{-1}) × (6.022 × 10^{23} mol^{-1}) = 7.2 × 10^{23}

Answer: **7.2 × 10^{23}**

Marks
4

- Lead ions react with bromide ions according to the following equation.

$$Pb^{2+}(aq) + 2Br^{-}(aq) \rightarrow PbBr_2(s)$$

If 0.040 M lead(II) nitrate solution (100.0 mL) is added to 0.020 M potassium bromide solution (300.0 mL), what amount (in mol) of lead(II) bromide precipitates?

The number of moles of Pb^{2+} ions in 100.0 mL of a 0.040 M solution of $Pb(NO_3)_2$ is:

number of moles = concentration × volume
= (0.040 mol L^{-1}) × (0.1000 L) = 0.0040 mol

The number of moles of Br^- ions in 300.0 mL of a 0.020 M solution of KBr is:

number of moles = (0.020 mol L^{-1}) × (0.3000 L) = 0.0060 mol

The precipitation reaction requires 2 Br^- ions for every Pb^{2+} ion. As there is *less* than twice as much Br^- than Pb^{2+}, it is the Br^- that is the limiting reagent.

From the chemical equation, 2 mol of Br^- leads to 1 mol of $PbBr_2$(s) and so:

number of moles of $PbBr_2$(s) formed = ½ × 0.0060 mol = 0.0030 mol

Answer: **0.0030 mol**

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

NO_3^- is not involved in the reaction: it is a spectator ion. The amount present are the reaction is the same as at the beginning. When 1 mol of $Pb(NO_3)_2$ dissolves, it forms 1 mol of Pb^{2+}(aq) and 2 mol of NO_3^-(aq). Hence, the number of moles of NO_3^-(aq) present is:

number of moles = 2 × (0.040 mol L^{-1}) × (0.1000 L) = 0.0080 mol

After the solutions are mixed, the total volume is (100.0 + 300.0) mL = 400.0 mL.

This amount is now present in this volume and so has a concentration:

concentration = number of moles / volume = 0.0080 mol / 0.4000 L = 0.020 M

Answer: **0.020 M**

Marks
2

- An unknown liquid contains H: 5.90 % and O: 94.1 % by mass and has a molar mass of 33.9 g mol^{-1}. What is its molecular formula?

The liquid contains 5.90% H and so 94.1% O.

	H	**O**
percentage	**5.90**	**94.1**
divide by atomic mass	$\frac{5.90}{1.008} = 5.85$	$\frac{94.1}{16.00} = 5.88$
divide by smallest value	**1**	**1**

The ratio of H : O is 1 : 1 and so the empirical formula is HO.

The molecular formula is $(HO)_n$. The molar mass is:

molar mass = $n \times (1.008 + 16.00)$ g mol^{-1} = $17.008n$ g mol^{-1}

As the molar mass is 33.9 g mol^{-1}, $n = 2$ and the molecular formula is $(HO)_2$ or H_2O_2. It is hydrogen peroxide.

Answer: **H_2O_2**

Marks
5

- A solution is prepared by dissolving lead(II) nitrate (33.12 g) in 1.00 L of water. Write the balanced ionic equation for this dissolution reaction.

$Pb(NO_3)_2(s) \rightarrow Pb^{2+}(aq) + 2NO_3^-(aq)$

When a 100.0 mL portion of this solution is mixed with a solution of potassium iodide (0.300 M, 150.0 mL), a bright yellow precipitate of lead(II) iodide forms. Write the balanced ionic equation for this precipitation reaction.

$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$

What mass of lead(II) iodide is formed?

The formula mass of $Pb(NO_3)_2$ is:

formula mass = (207.2 (Pb) + 2 × 14.01 (N) + 6 × 16.00 (O)) g mol^{-1}
= 331.22 g mol^{-1}

The number of moles in 33.12 g is therefore:

$$\text{number of moles} = \frac{\text{mass}}{\text{formula mass}} = \frac{33.12\text{ g}}{331.22\text{ g mol}^{-1}} = 0.1000\text{ mol}$$

If this dissolved in 1.00 L and a 100.0 mL portion is taken, this will contain 0.01000 mol of Pb^{2+}(aq).

150.0 mL of a 0.300 M solution of KI contains:

number of moles = concentration × volume
= 0.300 mol L^{-1} × 0.1500 L = 0.0450 mol

The precipitation reaction requires 2 mol of I^-(aq) for every 1 mol of Pb^{2+}(aq). The 0.01000 mol of Pb^{2+}(aq) that is present requires 0.02000 mol of I^-(aq). As there is more I^-(aq) than this present, I^-(aq) is in excess and Pb^{2+}(aq) is the limiting reagent.

From the precipitation reaction, 1 mol of Pb^{2+}(aq) will produce 1 mol of PbI_2(s). Therefore 0.01000 mol of Pb^{2+}(aq) will produce 0.01000 mol of PbI_2(s).

The formula mass of PbI_2 is:

formula mass = (207.2 (Pb) + 2 × 126.9 (I)) g mol^{-1}
= 461.0 g mol^{-1}

The mass of 0.01000 mol is therefore:

mass = formula mass × number of moles = 461.0 g mol^{-1} × 0.01000 mol
= 4.61 g

Answer: **4.61 g**

ANSWER CONTINUES ON THE NEXT PAGE

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

As described above reaction of 0.01000 mol of Pb^{2+}(aq) requires 0.02000 mol of I^-(aq). As 0.0450 mol are initially present, there are (0.0450 – 0.02000) mol = 0.0250 mol of I^-(aq) after the precipitation reaction.

After mixing the two solutions, the total volume becomes (100.0 + 150.0) mL = 250.0 mL. The final concentration of I^-(aq) is therefore:

$$\textbf{concentration} = \frac{\textbf{number of moles}}{\textbf{volume}} = \frac{\textbf{0.0250 mol}}{\textbf{0.2500 L}} = \textbf{0.100 mol L}^{-1} = \textbf{0.100 M}$$

Answer: **0.100 M**

- Three different oxides of lead are known. The oxide that is red in colour is found to consist of 90.67 % lead. What is its empirical formula? **2**

The oxide contains 90.67 % Pb and so 9.33 % O.

	Pb	**O**
percentage	**90.67**	**9.33**
divide by atomic mass	$\frac{90.67}{207.2} = 0.438$	$\frac{9.33}{16.00} = 0.583$
divide by smallest value	**1**	**1.33**

The ratio of Pb : O is 1 : 1.33 or 3: 4. The empirical formula is thus Pb_3O_4.

Answer: **Pb_3O_4**

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:

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

The balanced equation for the combustion of butane is:

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

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:

$$\text{number of moles of } O_2 \text{ needed} = \frac{13}{2} \times 0.10\text{ mol} = 0.65\text{ mol}$$

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

- 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?

2

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

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

Answer: **$BaCO_3$**

Marks
6

- 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.

$Pb^{2+}(aq) + 2I^{-}(aq) \rightarrow PbI_2(s)$

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^- = $(0.080 \text{ mol L}^{-1}) \times (0.0500 \text{ L}) = 0.0040 \text{ mol}$

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 mol = 0.0020 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**

- What mass of calcium chloride is required to make 250 mL of a 0.1 M solution? **3**

The formula mass of calcium chloride, $CaCl_2$, is

formula mass = 40.08 (Ca) + 2 × 35.45 (Cl) = 110.98

The number of moles in the solution is given by:

$$\textbf{number of moles = concentration × volume} = 0.1 \times \frac{250}{1000} = 0.025 \text{ mol}$$

The mass required is therefore:

mass = number of moles × formula mass = (0.025) × (110.98) = 3 g

Answer: **3 g**

What amount of chloride ions (in mol) is present in 30.0 mL of this solution?

One moles of $CaCl_2$(s) dissolves to give two moles of Cl^-(aq) ions. Therefore, the number of moles present is:

$$\textbf{number of moles = concentration × volume} = (2 \times 0.1) \times \frac{30}{1000} = 0.006 \text{ mol}$$

Answer: **0.006 mol**

Marks
4

- The complete combustion of butane, C_4H_{10}, in air gives water and carbon dioxide as the products. Write a balanced equation for this reaction.

$2C_4H_{10}(g) + 13O_2(g) \rightarrow 8CO_2(g) + 10H_2O(g)$ or

$C_4H_{10}(g) + {}^{13}/_{2}O_2(g) \rightarrow 4CO_2(g) + 5H_2O(g)$

What mass of oxygen is required for the complete combustion of 454 g of butane and what masses of carbon dioxide and water are produced?

The molar mass of butane is (4 × 12.01 (C)) + (10 × 1.008 (H)) = 58.12. Therefore, the amount of butane in 454 g is:

$$\text{number of moles of } C_4H_{10} = \frac{\text{mass}}{\text{molar mass}} = \frac{454}{58.12} = 7.81\,\text{mol}$$

From the chemical equation, each mole of C_4H_{10} requires ${}^{13}/_{2}$ moles of O_2 and produces 4 moles of CO_2 and 5 moles of H_2O.

Therefore:

number of moles of O_2 = ${}^{13}/_{2}$ × 7.81 = 50.8 mol
number of moles of CO_2 = 4 × 7.81 = 31.2 mol
number of moles of O_2 = 5 × 7.81 = 39.1 mol

The molar masses of O_2, CO_2 and H_2O are:

molar mass of O_2 = 2 × 16.00 = 32.00
molar mass of CO_2 = 12.01 (C) + (2 × 16.00) = 44.01
molar mass of H_2O = (2 × 1.008 (H)) + 16.00 (O)= 18.016

Therefore:

mass of O_2 = number of moles × molar mass = 50.8 × 32.00 = 1620 g = 1.62 kg
mass of CO_2 = 31.2 × 44.01 = 1380 g = 1.38 kg
mass of H_2O = 39.1 × 18.016 = 704 g = 0.704 kg

- During physical activity, lactic acid forms in the muscle tissue and is responsible for muscle soreness. Elemental analysis shows that it contains by mass 40.0% C, 6.71% H and 53.3% O. Determine the empirical formula of lactic acid.

4

	C	H	O
amount in 100 g	**40.0**	**6.71**	**53.3**
ratio (divide by atomic mass)	$\frac{40.0}{12.01} = 3.33$	$\frac{6.71}{1.008} = 6.66$	$\frac{53.3}{16.00} = 3.33$
divide by smallest	$\frac{3.33}{3.33} = 1.00 \sim 1$	$\frac{6.66}{3.33} = 2.00 \sim 2$	$\frac{3.33}{3.33} = 1.00$

The simplest possible ratio of C:H:O is thus 1:2:1 and the empirical formula is CH_2O.

Answer: **CH_2O**

Given that lactic acid has a molar mass of 90.08 g mol^{-1}, determine its molecular formula.

The molecular formula is $(CH_2O)_n$ so the molar mass is:

molar mass = n × (12.01 (C) + 2 × 1.008 (H) + 16.00 (O))
= 30.026n = 90.08 so n = 3

The molecular formula is thus $(CH_2O)_3$ or $C_3H_6O_3$

Answer: **$C_3H_6O_3$**

Marks 4

- If 50 mL of a 0.10 M solution of $AgNO_3$ is mixed with 50 mL of a 0.40 M solution of Na_2CO_3, what mass of Ag_2CO_3 will precipitate from the reaction?

The ionic equation for the precipitation reaction is:

$$2Ag^+(aq) + CO_3^{2-}(aq) \rightarrow Ag_2CO_3(s)$$

Thus, two moles of Ag^+(aq) are required for every one mole of CO_3^{2-}(aq).

The number of moles of Ag^+(aq) and CO_3^{2-}(aq) are given by:

$$n(Ag^+(aq)) = \text{concentration} \times \text{volume} = 0.10 \times \frac{50}{1000} = 0.0050 \text{ mol}$$

$$n(CO_3^{2-}(aq)) = 0.40 \times \frac{50}{1000} = 0.020 \text{ mol}$$

There is insufficient Ag^+(aq) to react with all of the CO_3^{2-}(aq) and so it is Ag^+(aq) is the limiting reagent. From the chemical equation, 1 mole of Ag_2CO_3(s) is produced from every two moles of Ag^+(aq) ions. The amount of Ag_2CO_3(s) produced is therefore:

$$n(Ag_2CO_3(s)) = ½ \times n(Ag^+(aq)) = ½ \times 0.0050 = 0.0025 \text{ mol}$$

The formula mass of Ag_2CO_3 is (2 × 107.87 (Ag)) + 12.01 (C) + (3 × 16.00 (O)) = 275.75. This number of moles thus corresponds to a mass of:

$$\text{mass of } Ag_2CO_3 = \text{number of moles} \times \text{formula mass} = 0.0025 \times 275.75 = 0.69 \text{ g}$$

Answer: **0.69 g**

What is the final concentration of CO_3^{2-} ions in the solution after the above reaction?

From the chemical equation, one mole of Ag_2CO_3(s) is produced from every mole of CO_3^{2-} which reacts. Therefore 0.0025 mol of CO_3^{2-} reacts. This leaves:

$$\text{number of moles of unreacted } CO_3^{2-} = 0.020 - 0.0025 = 0.018 \text{ mol}$$

The total volume of the solution after mixing is (50 + 50) = 100 mL. The final concentration is therefore:

$$\text{concentration} = \frac{\text{number of moles}}{\text{volume}} = \frac{0.018}{100/1000} = 0.18 \text{ M}$$

Answer: **0.18 M**

ANSWER CONTINUES ON THE NEXT PAGE

• Give balanced ionic equations for the reactions that occur in each of the following cases.	**3**
Sodium metal is added to excess water. **$2Na(s) + 2H_2O(l) \rightarrow 2Na^+(aq) + 2OH^-(aq) + H_2(g)$**	
Solutions of cobalt(II) nitrate and sodium phosphate are mixed. **$3Co^{2+}(aq) + 2PO_4^{3-}(aq) \rightarrow Co_3(PO_4)_2(s)$**	
Solid calcium carbonate is dissolved in dilute nitric acid. **$CaCO_3(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + CO_2(g) + H_2O(l)$**	

Marks

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

${}^{234}_{90}\text{Th} \rightarrow \boxed{{}^{234}_{91}\textbf{Pa}} + {}^{0}_{-1}\text{e}$
${}^{234}_{92}\text{U} \rightarrow \boxed{{}^{230}_{90}\textbf{Th}} + {}^{4}_{2}\text{He}$

- A nugget contains 2.6×10^{24} atoms of gold. What amount of gold (in mol) is in this nugget and what is its mass (in kg)? **2**

One mole of gold corresponds to Avogadro's number, 6.022×10^{23}, atoms. 2.6×10^{24} atoms therefore corresponds to:

$$\textbf{number of moles} = \frac{\textbf{number of atoms}}{\textbf{Avogadro's number}} = \frac{2.6 \times 10^{24}}{6.022 \times 10^{23}} = \textbf{4.3 mol}$$

As one mole of gold has a mass, corresponding to the atomic mass, of 196.97 g. 4.3 mol of gold therefore corresponds to:

mass = number of moles × atomic mass = 4.3 × 196.97 = 850 g = 0.85 kg

(Note that the number of atoms is given to 2 significant figures in the question and this is reflected in the answers).

Amount: **4.3 mol**	Mass: **0.85 kg**

Marks
7

- The complete combustion of propane, C_3H_8, in air gives water and carbon dioxide as the products? Write a balanced equation for this reaction.

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(l)$

What mass of oxygen is required for the complete combustion of 454 g of propane and what masses of CO_2 and H_2O are produced?

The molar mass of propane is (3 × 12.01 (C)) + (8 × 1.008 (H)) = 44.094. Therefore, 454 g corresponds to:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{454}{44.094} = 10.3 \text{ mol}$$

5 mol of $O_2(g)$ is required for every 1 mol of propane. Therefore, 5 × 10.3 = 51.5 mol of O_2 is required. The molar mass of O_2 is (2 × 16.00) = 32.00 so the mass of O_2 required is:

mass = number of moles × molar mass = 51.5 × 32.00 = 1650 g = 1.65 kg

3 mol of CO_2 and 4 mol of H_2O are produced for every 1 mol of propane. Therefore, 3 × 10.3 = 30.9 mol of CO_2 and 4 × 10.3 = 41.2 mol of H_2O are produced.

The molar mass of CO_2 is (12.01 (C)) + (2 × 16.00 (O)) = 44.01 and the molar mass of H_2O is (2 × 1.008 (H)) + (16.00 (O)) = 18.016.

The masses of CO_2 and H_2O produced are therefore:

mass of CO_2 = 30.9 × 44.01 = 1360 g = 1.36 kg
mass of H_2O = 41.2 × 18.016 = 742 g = 0.742 kg

(Note that the mass of propane is given to three significant figures in the question and this is reflected in each answer).

Explain the "law of conservation of mass". Show whether or not the above combustion conforms to this law.

The law of conservation of mass states that mass may neither be created nor destroyed. In this reaction:

mass of reactants = 0.454 kg (C_3H_8) + 1.65 kg (O_2) = 2.10 kg
mass of products = 1.36 kg (CO_2) + 0.742 kg (H_2O) = 2.10 kg

This combustion obeys the law (to 3 significant figures).

Marks
3

- The reaction of methane and water is one way to prepare hydrogen for use as a fuel.

$$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3H_2(g)$$

Which compound is the limiting reactant if you begin with 995 g of methane and 2510 g of water?

The molar mass of methane, CH_4, is (12.01 (C)) + (4 × 1.008 (H)) = 16.042. The number of moles of methane is therefore:

$$\textbf{moles of methane} = \frac{\textbf{mass}}{\textbf{molar mass}} = \frac{\textbf{995}}{\textbf{16.042}} = \textbf{62.0 mol}$$

The molar mass of water, H_2O, is (2 × 1.008 (H)) + (16.00 (H)) = 18.016. The number of moles of water is therefore:

$$\textbf{moles of methane} = \frac{\textbf{mass}}{\textbf{molar mass}} = \frac{\textbf{2510}}{\textbf{18.016}} = \textbf{139 mol}$$

As the reaction is a 1:1 reaction of methane and water, methane is the limiting reagent.

Answer: **methane, CH_4**

What mass of the excess reactant remains when the reaction is completed?

As the reaction is a 1:1 reaction, (139 – 62.0) = 77 mol of H_2O will be left unreacted. This corresponds to a mass of:

mass of water = moles of water × molar mass = 77 × 18.016 = 1400 g = 1.4 kg.

Answer: **1.4 kg**

Marks

4

- An unknown compound contains carbon and hydrogen only. If 0.0956 g of the compound is burned in oxygen, 0.300 g of CO_2 and 0.123 g of H_2O are isolated. What is the unknown compound's empirical formula?

The molar mass of CO_2 is (12.01 (C)) + (2 × 16.00 (O)) = 44.01. The molar mass of H_2O is (2 × 1.008 (H)) + (16.00 (O)) = 18.016. The number of moles of each after burning is therefore:

$$n_{CO_2} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.300}{44.01} = 0.00682\,\text{mol}, \quad n_{H_2O} = \frac{0.123}{18.016} = 0.00683\,\text{mol}$$

The moles of C in the compound is equal to the number of moles of CO_2, as the latter possesses one carbon atom per molecular unit.

The moles of H in the compound is equal to 2 × number of moles of H_2O, as the latter contains two hydrogen atoms per molecular unit.

The C:H ratio is therefore 1:2

Answer: **CH_2**

If its molar mass is found to be 70.1 g mol^{-1}, what is its molecular formula?

If the molar mass = 70.1, the number of moles in 0.0956 g is:

$$\text{number of moles} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.0956}{70.1} = 0.00136\,\text{mol}$$

As 0.00136 mol contains 0.683 mol of carbon, 1 mol contains $\frac{0.00683}{0.00136} = 5.01\,\text{mol}$

As C:H is 1:2, the compound must contain 10 hydrogen atoms.

Answer: **C_5H_{10}**

1

- What amount (in mol) of chloride ion is contained in 100 mL of 0.25 M magnesium chloride solution?

Magnesium chloride dissolves according to the equation, $MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2Cl^-(aq)$ so that two moles of chloride is produced for every mole of $MgCl_2$ present. The number of moles of $MgCl_2$ present is :

$$\text{number of moles} = \text{concentration} \times \text{volume} = 0.25 \times \frac{100}{1000} = 0.025\ \text{mol}$$

The number of moles of $Cl^-(aq)$ is therefore 2 × 0.025 = 0.050 mol

Answer: **0.050 mol**

- If 25.0 mL of 1.50 M hydrochloric acid is diluted to 500 mL, what is the molar concentration of the diluted acid? **1**

The number of moles of HCl present in 25.0 mL of a 1.50 M solution is:

$$\textbf{number of moles} = \textbf{concentration} \times \textbf{volume} = 1.50 \times \frac{25}{1000} = 0.0375 \textbf{ mol}$$

This number of moles in a 500 mL solution gives a concentration of:

$$\textbf{concentration} = \frac{\textbf{number of moles}}{\textbf{volume}} = \frac{0.0375}{(500/1000)} = 0.0750 \textbf{ M}$$

Answer: **0.0750 M**

- Write a balanced **ionic** equation for the reaction of solid sodium hydrogencarbonate, $NaHCO_3$, and dilute sulfuric acid, H_2SO_4. **2**

$$\mathbf{NaHCO_3(s) + H^+(aq) \rightarrow Na^+(aq) + H_2O(l) + CO_2(g)}$$

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

$$\textbf{The percentage of boron} = \frac{\textbf{mass of boron in one mole of hydride}}{\textbf{molar mass of hydride}} \times 100\%$$

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

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

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

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

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

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 (or sulfur tetrafluoride)**
K_2CrO_4	potassium chromate

Marks

- Solid sodium hydroxide reacts with carbon dioxide to produce sodium carbonate and water. Calculate the mass of sodium hydroxide required to prepare 53.0 g of sodium carbonate. **3**

The chemical reaction is:

$$2NaOH + CO_2 \rightarrow Na_2CO_3 + H_2O$$

The molar mass of Na_2CO_3 is:

molar mass = (2 × 22.99 (Na)) + 12.01 (C) + (3 × 16.00 (O))) = 105.99 g mol^{-1}

The number of moles of this in 53 g is therefore:

$$\text{number of moles} = \frac{\text{mass (in g)}}{\text{molar mass (in g mol}^{-1})} = \frac{53.0}{105.99} \text{ mol} = 0.50 \text{ mol}$$

For every 1 mole of Na_2CO_3 produced, 2 moles of NaOH are required. To make 0.50 mol therefore requires 1.00 mol of NaOH. The molar mass of NaOH is:

molar mass = 22.99 (Na) + 16.00 (O) + 1.008 (H) = 39.998 g mol^{-1}

As 1.00 mol is required, the mass required is 40.0 g.

Answer: **40.0 g**

- Analysis of an unknown compound returned the following percentage composition by weight: **3**

  nitrogen: 26.2%; chlorine: 66.4% hydrogen 7.5%

  What is the empirical formula of this compound?

	N	**Cl**	**H**
amount in 100 g	**26.2**	**66.4**	**7.5**
ratio (divide by atomic mass)	$\frac{26.2}{14.01} = 1.87$	$\frac{66.4}{35.45} = 1.87$	$\frac{7.5}{1.08} = 7.44$
divide by smallest	$\frac{1.87}{1.87} = 1.00 \sim 1$	$\frac{1.87}{1.87} = 1.00 \sim 1$	$\frac{7.44}{1.87} = 3.98 \sim 4$

Answer: **$NClH_4$**

Marks

- The relative atomic mass of magnesium is reported as 24.3. Show how this figure is calculated given the natural abundances of the following isotopes of magnesium: ^{24}Mg (79.0 %); ^{25}Mg (10.0 %); ^{26}Mg (11.0 %). **2**

The relative atomic mass of magnesium is the weighted average of the masses of its isotopes:

$$\left(24\times\frac{79.0}{100}\right)+\left(25\times\frac{10.0}{100}\right)+\left(26\times\frac{11.0}{100}\right)=24.3\,\mathrm{g\,mol^{-1}}$$

- With examples, briefly explain what allotropes are. **2**

Allotropes are different structural arrangements of the same atoms of an element.

Carbon occurs naturally as either graphite, which consists of sheets of planar hexagonal rings, and diamond, a three dimensional structure with tetrahedrally coordinated carbon. carbon. Oxygen exists as either the gaseous diatomic O_2 molecule or the gaseous triatomic O_3 (ozone).

- Complete the following table. **2**

Formula	Name
Na_2CO_3	**sodium carbonate**
Fe_2O_3	iron(III) oxide
PCl_3	**phosphorus trichloride**
NH_3	ammonia