

- 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:

$$\text{molar mass} = (26.98 (\text{Al}) + 3 \times 19.00 (\text{F})) \text{ g mol}^{-1} = 83.98 \text{ g mol}^{-1}$$

The number of moles in 0.50 g is therefore:

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

The concentration of this amount in 800.0 mL is then:

$$\begin{aligned} \text{concentration} &= \text{number of moles} / \text{volume} \\ &= 0.0060 \text{ mol} / 0.8000 \text{ L} = 0.0074 \text{ mol L}^{-1} \end{aligned}$$

Answer: **0.0074 mol L⁻¹ or 0.0074 M**

What is $[\text{F}^-]$ in this solution?

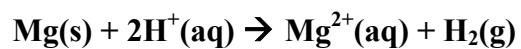
As the formula is AlF_3 , dissolution results in $3\text{F}^-(\text{aq})$ per formula unit.

$$[\text{F}^-(\text{aq})] = 3 \times 0.0074 \text{ mol L}^{-1} = 0.022 \text{ mol L}^{-1}$$

Answer: **0.022 mol L⁻¹ or 0.022 M**

THE REMAINDER OF THIS PAGE IS FOR ROUGH WORKING ONLY.

- 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.

Marks
3

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:

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

**From the chemical equation, each mol of Mg that reacts will give one mol of H₂.
Hence,**

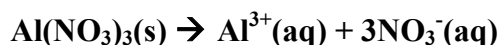
$$\text{number of moles of H}_2 = 0.21 \text{ mol.}$$

Answer: **0.21 mol**

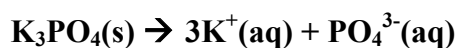
- A 0.060 M solution of aluminium nitrate and a 0.080 M solution of potassium phosphate are prepared by dissolving $\text{Al}(\text{NO}_3)_3$ and K_3PO_4 in water. Write the ionic equations for these two dissolutions reactions.

Marks
7

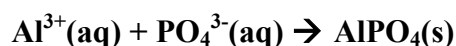
Dissolution
of $\text{Al}(\text{NO}_3)_3$



Dissolution
of K_3PO_4



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



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 $\text{Al}(\text{NO}_3)_3$ contains:

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

50.0 mL of a 0.080 M solution of K_3PO_4 contains:

$$\begin{aligned} \text{number of moles of } \text{PO}_4^{3-} &= \text{concentration} \times \text{volume} = c \times V \\ &= 0.080 \text{ mol L}^{-1} \times 0.0500 \text{ L} = 0.0040 \text{ mol} \end{aligned}$$

As the ionic equation has a 1 : 1 ratio of $\text{Al}^{3+} : \text{PO}_4^{3-}$ 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:

$$\text{number of moles of } \text{Al}^{3+} \text{ remaining} = (0.0060 - 0.0040) \text{ mol} = 0.0020 \text{ mol}$$

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

$$\begin{aligned} \text{concentration} &= \text{number of moles} / \text{volume} = n / V \\ &= 0.0020 \text{ mol} / 0.1500 \text{ L} = 0.013 \text{ mol L}^{-1} \end{aligned}$$

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?

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The mineral contains 84.80% U and 15.20% O.

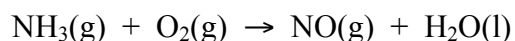
	U	O
percentage	84.80	15.20
divide by atomic mass	$\frac{84.80}{238.03} = 0.356$	$\frac{15.20}{16.00} = 0.950$
divide by smallest value	1	2.67

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.

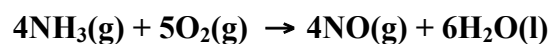
The empirical formula is U₃O₈.

Answer: U₃O₈

- Balance the following equation:



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

The molar mass of H_2O is:

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

Hence, the number of moles of water produced is:

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

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:

$$\text{number of moles of } \text{NH}_3 = (4/6) \times 7.771 \text{ mol} = 5.18 \text{ mol}$$

The molar mass of NH_3 is:

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

The mass of NH_3 in 5.18 mol is therefore:

$$\text{mass} = \text{number of moles} \times \text{molar mass} = (5.18 \text{ mol}) \times (17.034 \text{ g mol}^{-1}) = 88.2 \text{ g}$$

Answer: **88.2 g**

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

2

The volume of the block is:

$$V = \text{length} \times \text{width} \times \text{height} = (2.0 \times 2.0 \times 3.0) \text{ cm}^3 = 12 \text{ cm}^3$$

The mass can then be calculated from the density:

$$\text{density} = \text{mass} / \text{volume} \text{ or } \text{mass} = \text{density} \times \text{volume}$$

$$\text{mass} = (2.7 \text{ g cm}^{-3}) \times (12.0 \text{ cm}^3) = 32.4 \text{ g}$$

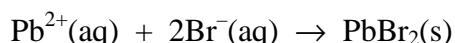
1 mol of Al has a mass equal to its atomic mass, 26.98 g mol^{-1} and contains $6.022 \times 10^{23} \text{ mol}^{-1}$. Hence, the number of atoms in 32.4 g is:

$$\begin{aligned} \text{number of atoms} &= \text{number of moles} \times \text{Avogadro's number} \\ &= (32.4 \text{ g} / 26.98 \text{ g mol}^{-1}) \times (6.022 \times 10^{23} \text{ mol}^{-1}) = 7.2 \times 10^{23} \end{aligned}$$

Answer: 7.2×10^{23}

Marks
4

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



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 $\text{Pb}(\text{NO}_3)_2$ is:

$$\begin{aligned}\text{number of moles} &= \text{concentration} \times \text{volume} \\ &= (0.040 \text{ mol L}^{-1}) \times (0.1000 \text{ L}) = 0.0040 \text{ mol}\end{aligned}$$

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

$$\text{number of moles} = (0.020 \text{ mol L}^{-1}) \times (0.3000 \text{ L}) = 0.0060 \text{ mol}$$

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 $\text{PbBr}_2(\text{s})$ and so:

$$\text{number of moles of } \text{PbBr}_2(\text{s}) \text{ formed} = \frac{1}{2} \times 0.0060 \text{ mol} = 0.0030 \text{ mol}$$

Answer: **0.0030 mol**

What is the final concentration of $\text{NO}_3^{-}(\text{aq})$ 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 $\text{Pb}(\text{NO}_3)_2$ dissolves, it forms 1 mol of $\text{Pb}^{2+}(\text{aq})$ and 2 mol of $\text{NO}_3^{-}(\text{aq})$. Hence, the number of moles of $\text{NO}_3^{-}(\text{aq})$ present is:

$$\text{number of moles} = 2 \times (0.040 \text{ mol L}^{-1}) \times (0.1000 \text{ L}) = 0.0080 \text{ mol}$$

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:

$$\text{concentration} = \text{number of moles} / \text{volume} = 0.0080 \text{ mol} / 0.4000 \text{ L} = 0.020 \text{ M}$$

Answer: **0.020 M**

- 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?

Marks
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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 $(\text{HO})_n$. The molar mass is:

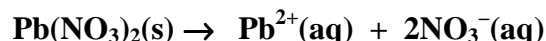
$$\text{molar mass} = n \times (1.008 + 16.00) \text{ g mol}^{-1} = 17.008n \text{ g mol}^{-1}$$

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

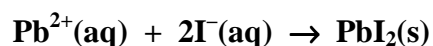
Answer: **H_2O_2**

- 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.

Marks
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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.



What mass of lead(II) iodide is formed?

The formula mass of $\text{Pb}(\text{NO}_3)_2$ is:

$$\begin{aligned} \text{formula mass} &= (207.2 \text{ (Pb)} + 2 \times 14.01 \text{ (N)} + 6 \times 16.00 \text{ (O)}) \text{ g mol}^{-1} \\ &= 331.22 \text{ g mol}^{-1} \end{aligned}$$

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 $\text{Pb}^{2+}(\text{aq})$.

150.0 mL of a 0.300 M solution of KI contains:

$$\begin{aligned} \text{number of moles} &= \text{concentration} \times \text{volume} \\ &= 0.300 \text{ mol L}^{-1} \times 0.1500 \text{ L} = 0.0450 \text{ mol} \end{aligned}$$

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

From the precipitation reaction, 1 mol of $\text{Pb}^{2+}(\text{aq})$ will produce 1 mol of $\text{PbI}_2(\text{s})$. Therefore 0.01000 mol of $\text{Pb}^{2+}(\text{aq})$ will produce 0.01000 mol of $\text{PbI}_2(\text{s})$.

The formula mass of PbI_2 is:

$$\begin{aligned} \text{formula mass} &= (207.2 \text{ (Pb)} + 2 \times 126.9 \text{ (I)}) \text{ g mol}^{-1} \\ &= 461.0 \text{ g mol}^{-1} \end{aligned}$$

The mass of 0.01000 mol is therefore:

$$\begin{aligned} \text{mass} &= \text{formula mass} \times \text{number of moles} = 461.0 \text{ g mol}^{-1} \times 0.01000 \text{ mol} \\ &= 4.61 \text{ g} \end{aligned}$$

Answer: **4.61 g**

ANSWER CONTINUES ON THE NEXT PAGE

What is the final concentration of $\Gamma(\text{aq})$ ions remaining in solution after the reaction is complete?

As described above reaction of 0.01000 mol of $\text{Pb}^{2+}(\text{aq})$ requires 0.02000 mol of $\Gamma(\text{aq})$. As 0.0450 mol are initially present, there are $(0.0450 - 0.02000)$ mol = 0.0250 mol of $\Gamma(\text{aq})$ 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 $\Gamma(\text{aq})$ is therefore:

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

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 Pb₃O₄.

Answer: **Pb₃O₄**

What mass of oxygen is required for the complete combustion of 5.8 g of butane, C₄H₁₀. How many moles of CO₂ and H₂O are produced?

Marks
4

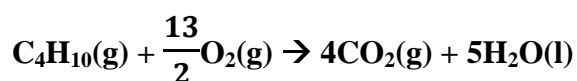
The molar mass of butane, C₄H₁₀, is:

$$\text{molar mass} = M = (4 \times 12.01 \text{ (C)} + 10 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 58.12 \text{ g mol}^{-1}$$

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:



Hence, 13/2 moles of O₂ are required and 4 moles of CO₂ and 5 moles of H₂O are produced for every mole of C₄H₁₀ which combusts. As 0.10 mol of C₄H₁₀ is present:

$$\text{number of moles of O}_2 \text{ needed} = \frac{13}{2} \times 0.10 \text{ mol} = 0.65 \text{ mol}$$

$$\text{number of moles of CO}_2 \text{ produced} = 4 \times 0.10 \text{ mol} = 0.40 \text{ mol}$$

$$\text{number of moles of H}_2\text{O produced} = 5 \times 0.10 \text{ mol} = 0.50 \text{ mol}$$

O₂ has a molar mass of (2 × 16.00) g mol⁻¹ = 32.00 g mol⁻¹. Hence the mass of O₂ required is:

$$\begin{aligned} \text{mass} &= \text{number of moles} \times \text{molar mass} \\ &= nM = (0.65 \text{ mol}) \times (32.00 \text{ g mol}^{-1}) = 21 \text{ g} \end{aligned}$$

- 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?

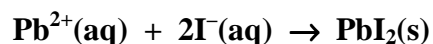
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	Ba	C	O
amount in 100 g	69.6	6.09	24.3
ratio (divide by atomic mass)	$\frac{69.6}{137.34} = 0.507$	$\frac{6.09}{12.01} = 0.507$	$\frac{24.3}{16.00} = 1.52$
divide by smallest	$\frac{0.507}{0.507} \sim 1$	$\frac{0.507}{0.507} \sim 1$	$\frac{1.52}{0.507}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

The simplest possible ratio of Ba:C:O is thus 1:1:3 and the empirical formula is BaCO₃.

Answer: **BaCO₃**

- 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.

Marks
6

What amount (in mol) of lead(II) iodide precipitates?

Before precipitation, the number of moles of $\text{Pb}^{2+}(\text{aq})$ and $\text{I}^{-}(\text{aq})$ present are:

$$\begin{aligned}\text{number of moles of } \text{Pb}^{2+} &= \text{concentration} \times \text{volume} \\ &= cV = (0.080 \text{ mol L}^{-1}) \times (0.1500 \text{ L}) = 0.012 \text{ mol} \\ \text{number of moles of } \text{I}^{-} &= (0.080 \text{ mol L}^{-1}) \times (0.0500 \text{ L}) = 0.0040 \text{ mol}\end{aligned}$$

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 $\text{PbI}_2(\text{s})$ is formed for every two moles of $\text{I}^{-}(\text{aq})$ present and hence:

$$\text{number of moles of } \text{PbI}_2(\text{s}) = \frac{1}{2} \times 0.0040 \text{ mol} = 0.0020 \text{ mol}$$

Answer: **0.0020 mol**

What amount (in mol) of $\text{Pb}^{2+}(\text{aq})$ ions remain in solution after the reaction?

From above, 0.012 mol of $\text{Pb}^{2+}(\text{aq})$ is initially present. The ionic equation shows that one mole of $\text{Pb}^{2+}(\text{aq})$ is lost for every one mole of $\text{PbI}_2(\text{s})$ formed. As 0.0020 mol of $\text{PbI}_2(\text{s})$ precipitates,

$$\text{number of moles of } \text{Pb}^{2+}(\text{aq}) \text{ left} = (0.012 - 0.0020) \text{ mol} = 0.010 \text{ mol}$$

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.

$\text{Pb}(\text{NO}_3)_2$ 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: **$0.12 \text{ mol L}^{-1} = 0.12 \text{ M}$**

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

3

The formula mass of calcium chloride, CaCl₂, is

$$\text{formula mass} = 40.08 (\text{Ca}) + 2 \times 35.45 (\text{Cl}) = 110.98$$

The number of moles in the solution is given by:

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

The mass required is therefore:

$$\text{mass} = \text{number of moles} \times \text{formula mass} = (0.025) \times (110.98) = 3 \text{ g}$$

Answer: **3 g**

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

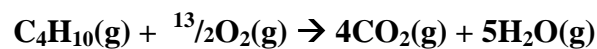
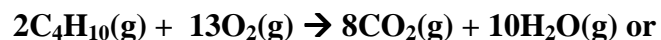
One moles of CaCl₂(s) dissolves to give two moles of Cl⁻(aq) ions. Therefore, the number of moles present is:

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

Answer: **0.006 mol**

- The complete combustion of butane, C₄H₁₀, in air gives water and carbon dioxide as the products. Write a balanced equation for this reaction.

Marks
4



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 $(4 \times 12.01 \text{ (C)}) + (10 \times 1.008 \text{ (H)}) = 58.12$. Therefore, the amount of butane in 454 g is:

$$\text{number of moles of C}_4\text{H}_{10} = \frac{\text{mass}}{\text{molar mass}} = \frac{454}{58.12} = 7.81 \text{ mol}$$

From the chemical equation, each mole of C₄H₁₀ requires $\frac{13}{2}$ moles of O₂ and produces 4 moles of CO₂ and 5 moles of H₂O.

Therefore:

$$\text{number of moles of O}_2 = \frac{13}{2} \times 7.81 = 50.8 \text{ mol}$$

$$\text{number of moles of CO}_2 = 4 \times 7.81 = 31.2 \text{ mol}$$

$$\text{number of moles of H}_2\text{O} = 5 \times 7.81 = 39.1 \text{ mol}$$

The molar masses of O₂, CO₂ and H₂O are:

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

$$\text{molar mass of CO}_2 = 12.01 \text{ (C)} + (2 \times 16.00) = 44.01$$

$$\text{molar mass of H}_2\text{O} = (2 \times 1.008 \text{ (H)}) + 16.00 \text{ (O)} = 18.016$$

Therefore:

$$\text{mass of O}_2 = \text{number of moles} \times \text{molar mass} = 50.8 \times 32.00 = 1620 \text{ g} = 1.62 \text{ kg}$$

$$\text{mass of CO}_2 = 31.2 \times 44.01 = 1380 \text{ g} = 1.38 \text{ kg}$$

$$\text{mass of H}_2\text{O} = 39.1 \times 18.016 = 704 \text{ g} = 0.704 \text{ kg}$$

- 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 CH₂O.

Answer: CH₂O

Given that lactic acid has a molar mass of 90.08 g mol⁻¹, determine its molecular formula.

The molecular formula is (CH₂O)_n so the molar mass is:

$$\begin{aligned} \text{molar mass} &= n \times (12.01 \text{ (C)} + 2 \times 1.008 \text{ (H)} + 16.00 \text{ (O)}) \\ &= 30.026n = 90.08 \text{ so } n = 3 \end{aligned}$$

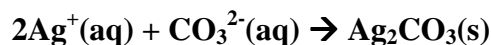
The molecular formula is thus (CH₂O)₃ or C₃H₆O₃

Answer: C₃H₆O₃

- 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?

Marks
4

The ionic equation for the precipitation reaction is:



Thus, two moles of $\text{Ag}^+(\text{aq})$ are required for every one mole of $\text{CO}_3^{2-}(\text{aq})$.

The number of moles of $\text{Ag}^+(\text{aq})$ and $\text{CO}_3^{2-}(\text{aq})$ are given by:

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

$$n(\text{CO}_3^{2-}(\text{aq})) = 0.40 \times \frac{50}{1000} = 0.020 \text{ mol}$$

There is insufficient $\text{Ag}^+(\text{aq})$ to react with all of the $\text{CO}_3^{2-}(\text{aq})$ and so it is $\text{Ag}^+(\text{aq})$ is the limiting reagent. From the chemical equation, 1 mole of $\text{Ag}_2\text{CO}_3(\text{s})$ is produced from every two moles of $\text{Ag}^+(\text{aq})$ ions. The amount of $\text{Ag}_2\text{CO}_3(\text{s})$ produced is therefore:

$$n(\text{Ag}_2\text{CO}_3(\text{s})) = \frac{1}{2} \times n(\text{Ag}^+(\text{aq})) = \frac{1}{2} \times 0.0050 = 0.0025 \text{ mol}$$

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

$$\text{mass of } \text{Ag}_2\text{CO}_3 = \text{number of moles} \times \text{formula mass} = 0.0025 \times 275.75 = 0.69 \text{ g}$$

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 $\text{Ag}_2\text{CO}_3(\text{s})$ is produced from every mole of CO_3^{2-} which reacts. Therefore 0.0025 mol of CO_3^{2-} reacts. This leaves:

$$\text{number of moles of unreacted } \text{CO}_3^{2-} = 0.020 - 0.0025 = 0.018 \text{ mol}$$

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

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

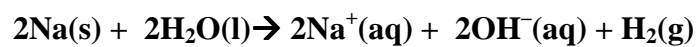
Answer: **0.18 M**

ANSWER CONTINUES ON THE NEXT PAGE

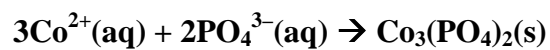
- Give balanced ionic equations for the reactions that occur in each of the following cases.

3

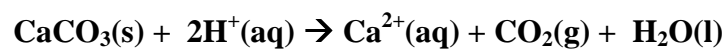
Sodium metal is added to excess water.



Solutions of cobalt(II) nitrate and sodium phosphate are mixed.

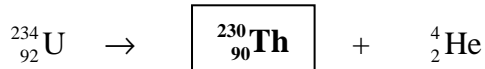
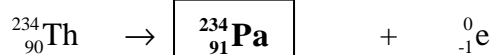


Solid calcium carbonate is dissolved in dilute nitric acid.



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

Marks
2



- 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:

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

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:

$$\text{mass} = \text{number of moles} \times \text{atomic mass} = 4.3 \times 196.97 = 850 \text{ g} = 0.85 \text{ kg}$$

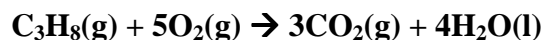
(Note that the number of atoms is given to 2 significant figures in the question and this is reflected in the answers).

Amount: **4.3 mol**

Mass: **0.85 kg**

- The complete combustion of propane, C₃H₈, in air gives water and carbon dioxide as the products? Write a balanced equation for this reaction.

Marks
7



What mass of oxygen is required for the complete combustion of 454 g of propane and what masses of CO₂ and H₂O are produced?

The molar mass of propane is $(3 \times 12.01 \text{ (C)}) + (8 \times 1.008 \text{ (H)}) = 44.094$. 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 O₂(g) is required for every 1 mol of propane. Therefore, $5 \times 10.3 = 51.5$ mol of O₂ is required. The molar mass of O₂ is $(2 \times 16.00) = 32.00$ so the mass of O₂ required is:

$$\text{mass} = \text{number of moles} \times \text{molar mass} = 51.5 \times 32.00 = 1650 \text{ g} = 1.65 \text{ kg}$$

3 mol of CO₂ and 4 mol of H₂O are produced for every 1 mol of propane. Therefore, $3 \times 10.3 = 30.9$ mol of CO₂ and $4 \times 10.3 = 41.2$ mol of H₂O are produced.

The molar mass of CO₂ is $(12.01 \text{ (C)}) + (2 \times 16.00 \text{ (O)}) = 44.01$ and the molar mass of H₂O is $(2 \times 1.008 \text{ (H)}) + (16.00 \text{ (O)}) = 18.016$.

The masses of CO₂ and H₂O produced are therefore:

$$\begin{aligned} \text{mass of CO}_2 &= 30.9 \times 44.01 = 1360 \text{ g} = 1.36 \text{ kg} \\ \text{mass of H}_2\text{O} &= 41.2 \times 18.016 = 742 \text{ g} = 0.742 \text{ kg} \end{aligned}$$

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

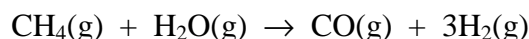
Explain the "law of conservation of mass". Show whether or not the above combustion conforms to this law.

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

$$\begin{aligned} \text{mass of reactants} &= 0.454 \text{ kg (C}_3\text{H}_8) + 1.65 \text{ kg (O}_2) = 2.10 \text{ kg} \\ \text{mass of products} &= 1.36 \text{ kg (CO}_2) + 0.742 \text{ kg (H}_2\text{O)} = 2.10 \text{ kg} \end{aligned}$$

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

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



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

Marks
3

The molar mass of methane, CH₄, is (12.01 (C)) + (4 × 1.008 (H)) = 16.042. The number of moles of methane is therefore:

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

The molar mass of water, H₂O, is (2 × 1.008 (H)) + (16.00 (O)) = 18.016. The number of moles of water is therefore:

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

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

Answer: **methane, CH₄**

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 H₂O will be left unreacted. This corresponds to a mass of:

$$\text{mass of water} = \text{moles of water} \times \text{molar mass} = 77 \times 18.016 = 1400 \text{ g} = 1.4 \text{ kg.}$$

Answer: **1.4 kg**

- An unknown compound contains carbon and hydrogen only. If 0.0956 g of the compound is burned in oxygen, 0.300 g of CO₂ and 0.123 g of H₂O are isolated. What is the unknown compound's empirical formula?

Marks
4

The molar mass of CO₂ is (12.01 (C)) + (2 × 16.00 (O)) = 44.01. The molar mass of H₂O is (2 × 1.008 (H)) + (16.00 (O)) = 18.016. The number of moles of each after burning is therefore:

$$n_{\text{CO}_2} = \frac{\text{mass}}{\text{molar mass}} = \frac{0.300}{44.01} = 0.00682 \text{ mol}, \quad n_{\text{H}_2\text{O}} = \frac{0.123}{18.016} = 0.00683 \text{ mol}$$

The moles of C in the compound is equal to the number of moles of CO₂, 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 H₂O, as the latter contains two hydrogen atoms per molecular unit.

The C:H ratio is therefore 1:2

Answer: CH₂

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: C₅H₁₀

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

1

Magnesium chloride dissolves according to the equation, MgCl₂(s) → Mg²⁺(aq) + 2Cl⁻(aq) so that two moles of chloride is produced for every mole of MgCl₂ present. The number of moles of MgCl₂ 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

1

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

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

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

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

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

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

- The element boron forms a series of hydrides, which includes B₂H₆, B₄H₁₀, B₅H₉, B₆H₁₀ and B₁₀H₁₄. Which one of these hydrides consists of 85.63% boron by mass?

Marks
2

The molar mass of the boranes are:

$$\text{molar mass of B}_2\text{H}_6 = (2 \times 10.81 \text{ (B)}) + (6 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 27.668 \text{ g mol}^{-1}$$

$$\text{molar mass of B}_4\text{H}_{10} = (4 \times 10.81 \text{ (B)}) + (10 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 53.32 \text{ g mol}^{-1}$$

$$\text{molar mass of B}_5\text{H}_9 = (5 \times 10.81 \text{ (B)}) + (9 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 63.122 \text{ g mol}^{-1}$$

$$\text{molar mass of B}_6\text{H}_{10} = (6 \times 10.81 \text{ (B)}) + (10 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 74.94 \text{ g mol}^{-1}$$

$$\text{molar mass of B}_{10}\text{H}_{14} = (10 \times 10.81 \text{ (B)}) + (14 \times 1.008 \text{ (H)}) \text{ g mol}^{-1} = 122.212 \text{ g mol}^{-1}$$

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

$$\text{percentage boron in B}_2\text{H}_6 = \frac{2 \times 10.81}{27.668} \times 100\% = 78.14\%$$

$$\text{percentage boron in B}_4\text{H}_{10} = \frac{4 \times 10.81}{53.32} \times 100\% = 81.10\%$$

$$\text{percentage boron in B}_5\text{H}_9 = \frac{5 \times 10.81}{63.122} \times 100\% = 85.63\%$$

$$\text{percentage boron in B}_6\text{H}_{10} = \frac{6 \times 10.81}{74.94} \times 100\% = 86.55\%$$

$$\text{percentage boron in B}_{10}\text{H}_{14} = \frac{10 \times 10.81}{122.12} \times 100\% = 88.45\%$$

Answer: B₅H₉

- Complete the following table.

Formula	Name
K ₂ SO ₄	potassium sulfate
CuCl ₂	copper(II) chloride
SF ₄	sulfur(IV) fluoride (or sulfur tetrafluoride)
K ₂ CrO ₄	potassium chromate

2

- 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.

Marks
3

The chemical reaction is:



The molar mass of Na_2CO_3 is:

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

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

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

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:

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

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:

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

What is the empirical formula of this compound?

3

	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

- 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 %).

Marks
2

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

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

- 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