Atoms, Bonds and Groups
Moles and Equations

108 Marks

1. (a) A student carries out a titration to find the concentration of some sulfuric acid.

The student finds that 25.00 cm^3 of 0.0880 mol dm^{-3} aqueous sodium hydroxide, NaOH, is neutralised by 17.60 cm^3 of dilute sulfuric acid, H_2SO_4.

$$H_2SO_4(aq) + 2NaOH(aq) \rightarrow Na_2SO_4(aq) + 2H_2O(l)$$

(i) Calculate the amount, in moles, of NaOH used.

NaOH: 25.00 cm^3 of 0.0880 mol dm^{-3}

$n(NaOH) = \frac{25.00}{1000} \times 0.0880 = 2.20 \times 10^{-3}$ mol

answer = ... 2.20×10^{-3} ..... mol

[1]

(ii) Determine the amount, in moles, of H_2SO_4 used.

From equation: $H_2SO_4 + 2\,NaOH \rightarrow Na_2SO_4 + 2H_2O$

1 : 2

$n(H_2SO_4) = \frac{2.20 \times 10^{-3} \text{ mol}}{2}$

answer = ... 1.10×10^{-3} ..... mol

[1]

(iii) Calculate the concentration, in mol dm^{-3}, of the sulfuric acid.

H_2SO_4 : 1.10×10^{-3} mol in 17.60 cm^3 = $\frac{17.60\ cm^3}{1000} = 0.0176\ dm^3$

$[H_2SO_4] = \frac{1.10 \times 10^{-3} \text{ mol}}{0.0176\ dm^3}$

answer = ... 6.25×10^{-2} ..... mol dm^{-3}

[1]

(b) After carrying out the titration in (a), the student left the resulting solution to crystallise. White crystals were formed, with a formula of $Na_2SO_4 \bullet \boldsymbol{x}\ H_2O$ and a molar mass of 322.1 g mol^{-1}.

(i) What term is given to the '$\bullet \boldsymbol{x}\ H_2O$' part of the formula?

Water of crystallisation .....

[1]

$M_r\ (Na_2SO_4) = (23\times2) + 32.1 + (16\times4)$
$= 142.1\ g\ mol^{-1}$

(ii) Using the molar mass of the crystals, calculate the value of ***x***.

142.1
$Na_2SO_4 \cdot x H_2O$
322.1

Mass of H_2O = 322.1 − 142.1
= 180g
$n(H_2O) = \frac{180g}{18} = 10\ mol$
$x = 10$

answer = ... 10 .....

[2]

[Total 6 marks]

2. The element strontium forms a nitrate, $Sr(NO_3)_2$, which decomposes on heating as shown below.

NO_3^- (O^{2-}, O)

$$2Sr(NO_3)_2(s) \rightarrow 2SrO(s) + 4NO_2(g) + O_2(g)$$

(N^{5+}, N^{4+})

(i) Using oxidation numbers, explain why the reaction involves both oxidation and reduction.

O goes from −2 to 0 - oxidised
N " " +5 to +4 - reduced

N is reduced and O is oxidised

[3]

(ii) A student heats 5.29 g of $Sr(NO_3)_2$ and collects the gas at room temperature and pressure, RTP.

$$2Sr(NO_3)_2(s) \rightarrow 2SrO(s) + 4NO_2(g) + O_2(g)$$
2 4 1

Calculate the volume of gas, in dm^3, obtained by the student at RTP.

Molar mass of $Sr(NO_3)_2$ = 211.6 g mol $^{-1}$.

Mass $Sr(NO_3)_2$ = 5.29g

$n(Sr(NO_3)_2) = \frac{5.29}{211.6} = 0.025\ mol$

2 mol $Sr(NO_3)_2$ → 4 mol NO_2
1 mol O_2
5 mol gas

answer = ... 1.5 ..... dm^3

[3]

[Total 6 marks]

0.025 mol $Sr(NO_3)_2$ → $\frac{5}{2} \times 0.025 = 0.0625$ mol

1 mole gas = 24 dm^3 RTP

Volume of gas = 24 × 0.0625
= 1.5 dm^3

3. In the sixteenth century, a large deposit of graphite was discovered in the Lake District.

People at the time thought that the graphite was a form of lead.

Nowadays, graphite is used in pencils but it is still referred to as 'pencil lead'.

A student decided to investigate the number of carbon atoms in a 'pencil lead'. He found that the mass of the 'pencil lead' was 0.321 g.

(i) Calculate the amount, in mol, of carbon atoms in the student's pencil lead.

Assume that the 'pencil lead' is pure graphite.

moles of C: $n(C) = \frac{\text{mass}}{A_r} = \frac{0.321}{12} = 0.02675$ moles

answer = ... 0.02675 ..... mol

[1]

(ii) Using the Avogadro constant, N_A, calculate the number of carbon atoms in the student's 'pencil lead'.

Avogadro N°: $= 6.02 \times 10^{23}$ - N° of atoms in 1 mole.

No of C atoms = no moles C $\times N_A$

$= 0.02675 \times 6.02 \times 10^{23} = 1.61 \times 10^{22}$

number of carbon atoms = ... 1.61×10^{22} .....

[1]

[Total 2 marks]

Ba^{2+}

4. The Group 2 element barium, Ba, is silvery white when pure but blackens when exposed to air.

The blackening is due to the formation of both barium oxide and barium nitride. The nitride ion is N^{3-}.

(a) Predict the formula of:

barium oxide ..... BaO ..... barium nitride ..... Ba_3N_2 .....

Ba^{2+} O^{2-} Ba^{2+} N^{3-}

[2]

(b) A 0.11 g sample of pure barium was added to 100 cm^3 of water.

$Ba(s) + 2H_2O(l) \rightarrow Ba(OH)_2(aq) + H_2(g)$

(i) Show that 8.0×10^{-4} mol of Ba were added to the water.

$n(Ba) = \frac{\text{mass}}{A_r} = \frac{0.11}{137.3}$

$= 8.0 \times 10^{-4}$ mol

[1]

1 mol Ba ⟶ 1 mol H_2

(ii) Calculate the volume of hydrogen, in cm^3, produced at room temperature and pressure.

$n(H_2) = n(Ba) = 8.0 \times 10^{-4}$ mol

$V = 24 \times 8.0 \times 10^{-4}$ mol $= 0.0192\ dm^3$

$1\ dm^3 = 1000\ cm^3$

$0.0192\ dm^3 = 19.2\ cm^3$

volume = 19.2 cm^3

[1]

(iii) Calculate the concentration, in mol dm^{-3}, of the $Ba(OH)_2(aq)$ solution formed.

$n(Ba(OH)_2) = n(Ba) = 8.0 \times 10^{-4}$ mol in 100 $cm^3 = \frac{100}{1000}\ dm^3$

$[Ba(OH)_2] = \frac{\text{moles}}{\text{vol}} = \frac{8.0 \times 10^{-4}}{0.1} = 8.0 \times 10^{-3}$ mol dm^{-3} $= 0.1\ dm^3$

concentration = 8.0×10^{-3} mol dm^{-3}

[1]

alkaline

(iv) State the approximate pH of the $Ba(OH)_2(aq)$ solution.

pH > 7 and < 15

[1]

(c) A student repeated the experiment in (b) using a 0.11 g sample of barium that had blackened following exposure to the air.

Suggest why the volume of hydrogen produced would be slightly less than the volume collected using pure barium.

less volume of Ba because some has reacted already to form BaO/Ba_3N_2

[1]

(d) Describe and explain the trend, down the group, in the reactivity of the Group 2 elements with water.

Reactivity increases down the group
Atomic radii increases with extra electron shells
This causes increased shielding.
Outer electrons are further from the nucleus
Thus nuclear attraction decreases
Therefore easier to remove an electron from outer shell
Ionisation energy decreases
Reactivity increases.

[5]

[Total 12 marks]

5. Calcium carbonate, $CaCO_3$, reacts with hydrochloric acid as shown in the equation below.

1 : 2

$$CaCO_3(s) + 2HCl\,(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

7.50×10^{-3} mol $CaCO_3$ reacts with 0.200 mol dm^{-3} HCl.

(i) Calculate the volume, in cm^3, of 0.200 mol dm^{-3} HCl required to react with 7.50×10^{-3} mol $CaCO_3$.

$n(HCl) = 7.50 \times 10^{-3} \times 2 = 0.015$ mol

$Vol = \frac{0.015 \text{ mol}}{0.200 \text{ mol dm}^{-3}}$

$= 0.075 \text{ dm}^3$

$= 0.075 \text{ dm}^3 \times 1000$

$= 75 \text{ cm}^3$

answer = 75 cm^3

[2]

(ii) Calculate the volume, in cm^3, of CO_2 formed at room temperature and pressure.

$n(CO_2) = n(CaCO_3) = 7.50 \times 10^{-3}$

Volume $CO_2 = 7.50 \times 10^{-3} \times 24$

$= 1.8\ dm^3$

$1\ dm^3 = 1000\ cm^3$ answer = ... 180 ..... cm^3

[1]

Volume $= 1.8 \times 1000$

$= 180\ cm^3$

[Total 3 marks]

6. Calcium and its compounds, have properties typical of Group 2 in the Periodic Table.

Calcium carbonate, $CaCO_3$, reacts with acids such as nitric acid.

A student neutralised 2.68 g of $CaCO_3$ with 2.50 mol dm^{-3} nitric acid, HNO_3.

The equation for this reaction is shown below.

$$CaCO_3(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + CO_2(g) + H_2O(l)$$

(i) Determine the amount, in mol, of $CaCO_3$ reacted.

$M_r(CaCO_3) = 40.1 + 12 + (16 \times 3) = 100.1\ g\,mol^{-1}$

$n(CaCO_3) = \frac{2.68\,g}{100.1\ g\,mol^{-1}}$

$= 0.0268\ mol$

amount = ... 0.0268 ..... mol

[2]

(ii) Calculate the volume, in cm^3, of CO_2 produced at room temperature and pressure.

$n(CO_2) = n(CaCO_3) = 0.0268\ mol$

Volume $= 0.0268 \times 24$

$= 0.6426\ dm^3$

$= 642.6\ cm^3$ volume = ... 643 ..... cm_3

[1]

$CaCO_3 + 2HNO_3$

1 : 2

(iii) Calculate the volume of 2.50 mol dm^{-3} HNO_3 needed to neutralise 2.68 g of $CaCO_3$.

$n(HNO_3) = 2n(CaCO_3) = 0.0268\ mol \times 2 = 0.0536\ mol$

$Volume = \frac{moles}{concentration} = \frac{0.0536\ mol}{2.5\ mol\ dm^{-3}} = 0.02144\ dm^3$

$= 21.44\ cm^3$

volume = 21.44 cm^3

[2]

[Total 5 marks]

7. Rubidium forms an ionic compound with silver and iodine. This compound has a potential use in miniaturised batteries because of its high electrical conductivity.

The empirical formula of this ionic compound can be calculated from its percentage composition by mass: Rb, 7.42%; Ag, 37.48%; I, 55.10%.

(i) Define the term *empirical formula*.

Simplest whole number ratio of elements.

[1]

(ii) Calculate the empirical formula of the compound.

	Rb	Ag	I
Mass %	7.42%	37.48%	55.1%
Mass in 100g:	7.42g	37.48g	55.1g
N° moles	$\frac{7.42g}{85.5}$	$\frac{37.48g}{107.9}$	$\frac{55.1g}{126.9}$
	= 0.08678	0.3474	0.4342
÷ By smallest:	$\frac{0.08678}{0.08678}$	$\frac{0.3474}{0.08678}$	$\frac{0.4342}{0.08678}$
=	1 :	4 :	5

Empirical Formula = $RbAg_4I_5$

[2]

[Total 3 marks]

8. Chlorine can be prepared by reacting concentrated hydrochloric acid with manganese(IV) oxide.

$$4HCl(aq) + MnO_2(s) \rightarrow Cl_2(g) + MnCl_2(aq) + 2H_2O(l)$$

(a) A student reacted 50.0 cm^3 of 12.0 mol dm^{-3} hydrochloric acid with an excess of manganese(IV) oxide.

(i) Calculate how many moles of HCl were reacted.

$n(HCl) = \frac{50.0}{1000} \times 12.0$ mol dm^{-3}

$= 0.6$ mol

answer = ... 0.6 ..... mol

[1]

(ii) Calculate the volume of $Cl_2(g)$ produced, in dm^3.

Under the experimental conditions, one mole of $Cl_2(g)$ occupies 24.0 dm^3.

$4HCl \rightarrow Cl_2$

$n(Cl_2) = \frac{0.6}{4} = 0.15$

Volume $= 0.15 \times 24$

$= 3.6$ dm^3

answer = ... 3.6 ..... dm^3

[2]

(b) In this reaction, chlorine is oxidised.

Use oxidation numbers to determine what is reduced.

.....

$4HCl(aq) + MnO_2(s) \rightarrow Cl_2(g) + MnCl_2(aq) + 2H_2O(l)$

Mn^{4+} Mn^{2+}

Mn has been reduced from +4 to +2

[2]

[Total 5 marks]

9. This question looks at the reaction of sodium with water and with oxygen.

A chemist reacted 0.0500 mol of sodium with water to form 50.0 cm^3 of aqueous sodium hydroxide.

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

A_r (Na) = 23 g mol^{-1}

(i) What mass of Na was reacted?

n(Na) = 0·05 mol

Mass = 0·05 mol × 23 g mol^{-1}

= 1·15 g

mass = 1·15 g

[1]

(ii) Calculate the volume of H_2, in dm^3, that would be produced at room temperature and pressure, r.t.p.

1 mol of gas molecules occupies 24.0 dm^3 at r.t.p.

2 Na → H_2

$n(H_2) = \frac{0.05}{2} = 0.025$ mol

Vol = 0·025 × 24 = 0·6 dm^3

volume = 0·6 dm^3

[2]

(iii) Calculate the concentration, in mol dm^{-3}, of NaOH that was formed.

n(NaOH) = n(Na) = 0·05 mol

Volume = 50 cm^3 = 0·05 dm^3

$[NaOH] = \frac{0.05 \text{ mol}}{0.05 \text{ dm}^3} = 1 \text{ mol dm}^{-3}$

concentration = 1·00 mol dm^{-3}

[1]

[Total 4 marks]

10. Sodium reacts with excess oxygen to form sodium peroxide, Na_2O_2.

Na_2O_2 is used in laundry bleaches. When added to water a reaction takes place forming an alkaline solution and hydrogen peroxide, H_2O_2.

(i) Construct a balanced equation for the formation of sodium peroxide from sodium.

2 Na + O_2 → Na_2O_2

[1]

(ii) Construct a balanced equation for the reaction of sodium peroxide with water.

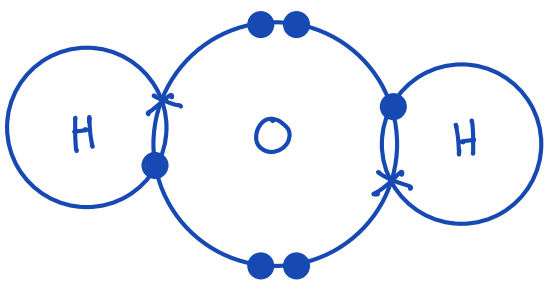
Na_2O_2 + 2 H_2O → H_2O_2 + 2 NaOH

[1]

(iii) Draw a 'dot-and-cross' diagram for a molecule of H_2O_2. Show outer electrons only.

[2]

[Total 4 marks]



$M_r(BaO) = 137.3 + 16$
$= 153.3 \text{ g mol}^{-1}$

11. Barium metal can be extracted from barium oxide, BaO, by reduction with aluminium.

$$6BaO + 2Al \rightarrow 3Ba + Ba_3Al_2O_6$$
6 : 3

Calculate the mass of barium metal that could be produced from reduction of 500 g of barium oxide using this method.

Mass BaO = 500 g $n(BaO) = \frac{500 \text{ g}}{153.3 \text{ g mol}^{-1}}$ answer = ... 223.9 ..... g

[Total 4 marks]

$n(BaO) = 3.2616 \text{ mol}$

$n(Ba) = 3.2616 \text{ mol} \times \frac{3}{6}$

$= 1.6308 \text{ mol}$ Mass $= 1.6308 \text{ mol} \times 137.3 \text{ g mol}^{-1} = 223.9 \text{ g}$

12. Water, ammonia and sulphur dioxide react together to form a compound **A** which has the following percentage composition by mass:

N, 24.12%;

H, 6.94%;

S, 27.61%;

O, 41.33%.

	N	H	S	O
Mass (g)	24.12	6.94	27.61	41.33
Moles	$\frac{24.12}{14}$	$\frac{6.94}{1}$	$\frac{27.61}{32.1}$	$\frac{41.33}{16}$
=	1.7229	6.94	0.8601	2.5831
÷ Smallest :	$\frac{1.7229}{0.8601}$	$\frac{6.94}{0.8601}$	$\frac{0.8601}{0.8601}$	$\frac{2.5831}{0.8601}$
	2 :	8 :	1 :	3

(i) Calculate the empirical formula of compound **A.**

$N_2H_8SO_3$

[2]

(ii) Suggest a balanced equation for the formation of compound **A** from the reaction of water, ammonia and sulphur dioxide.

$H_2O + 2\ NH_3 + SO_2 \rightarrow (NH_4)_2SO_3$

[1]

[Total 3 marks]

13. A student carried out two experiments using chlorine gas, $Cl_2(g)$.

(a) In the first experiment, the student bubbled chlorine through 120 cm^3 of an aqueous solution of 0.275 mol dm^{-3} sodium hydroxide, NaOH(aq).

The equation for this reaction is shown below.

$$Cl_2(g) + 2NaOH(aq) \rightarrow NaCl(aq) + NaClO(aq) + H_2O(l)$$

Under the reaction conditions, 1 mole of $Cl_2(g)$ occupies 24.0 dm^3.

(i) What is meant by the term *the mole*?

Amount of substance that has the same number of particles as there are atoms in 12g of ^{12}C (6×10^{23} Avogadro's Number)

[1]

(ii) How many moles of NaOH were in the 120 cm^3 volume of NaOH(aq)?

$n(NaOH) = \frac{120}{1000} \times 0.275 = 0.033$ mol

answer 0.033 mol

[1]

$1Cl_2 : 2NaOH$

(iii) Calculate the volume of $Cl_2(g)$ that was needed to react with the NaOH(aq) used.

$n(Cl_2) = \frac{n(NaOH)}{2} = \frac{0.033}{2} = 0.0165$ mol

answer 396 cm^3

$Vol = 0.0165 \times 24 = 0.396\ dm^3 = 396\ cm^3$

[2]

(iv) What is a common use for the solution that the student prepared?

Bleach

[1]

(b) In the second experiment, the student repeated the procedure in (a) but with hot concentrated sodium hydroxide. A different reaction took place in which sodium chlorate (V) was formed instead of NaClO.

Suggest the formula of sodium chlorate (V).

$NaClO_3$

[1]

[Total 6 marks]

14. Antimony is found naturally in a number of minerals including stibnite. Stibnite typically contains 5% of Sb_2S_3. Antimony can be obtained by reducing Sb_2S_3 with scrap iron.

$$Sb_2S_3 + 3Fe \rightarrow 2Sb + 3FeS$$

(i) How many moles of Sb_2S_3 are in 500 kg of a typical sample of stibnite containing 5% by mass of Sb_2S_3?

molar mass of Sb_2S_3 = 340 g mol^{-1}; relative atomic mass of Sb = 122

Mass $Sb_2S_3 = 500 \times \frac{5}{100} = 25$ kg

$= 2500$ g

$n(Sb_2S_3) = \frac{2500\ g}{340\ g\ mol^{-1}} = 73.53$

74 mol

[2]

(ii) Calculate the mass of antimony that could be obtained by processing 500 kg of stibnite.

$1\ Sb_2S_3 \longrightarrow 2\ Sb$

$n(Sb) = n(Sb_2S_3) \times 2$

$= 74 \times 2 = 147$ mol

Mass Sb $= 147 \times 122$

$= 17941$ g $= 17.94$ kg

mass = ... 17.9 ..... kg

[2]

[Total 4 marks]

15. Calcium oxide neutralises acids such as nitric acid. A student neutralised 1.50 g of CaO with 2.50 mol dm^{-3} nitric acid, HNO_3. The equation for this reaction is shown below.

$M_r(CaO) = 40.1 + 16$

$= 56.1$ g mol^{-1}

$CaO(s) + 2HNO_3(aq) \rightarrow Ca(NO_3)_2(aq) + H_2O(l)$

(i) How many moles of CaO were reacted?

$n(CaO) = \frac{1.50}{56.1} = 0.0267$

..... 0.0267 ..... mol

[2]

(ii) Calculate the volume of 2.50 mol dm^{-3} HNO_3 needed to exactly neutralise 1.50 g of CaO.

$1\ CaO = 2\ HNO_3$

$n(HNO_3) = 2 \times 0.0267 = 0.05348$ mol

$Vol = \frac{0.05348 \text{ mol}}{2.50 \text{ mol dm}^{-3}} = 0.02139\ dm^3 = 21.4\ cm^3$

volume = ... 21.4 ..... cm^3

[2]

[Total 4 marks]

16. The nitrate ion, NO_3^-, in $Ca(NO_3)_2$ contains both covalent and dative covalent bonds.

(i) What is the difference between a covalent bond and a dative covalent bond?

Dative covalent: bonded pair comes from same atom

Whereas a covalent bond is formed from sharing electrons from both atoms

[1]

(ii) Calcium nitrate decomposes on heating to form calcium oxide, oxygen and nitrogen(IV) oxide, NO_2.

Construct a balanced equation for this reaction.

$Ca(NO_3)_2 \longrightarrow CaO + 2\ NO_2 + \frac{1}{2}\ O_2$

[1]

[Total 2 marks]

17. Aqueous silver nitrate can be used as a test for halide ions. A student decided to carry out this test on a solution of magnesium chloride. The bottle of magnesium chloride that the student used showed the formula $MgCl_2.6H_2O$.

The student dissolved a small amount of $MgCl_2.6H_2O$ in water and added aqueous silver nitrate to the aqueous solution.

(i) What is the molar mass of $MgCl_2.6H_2O$?

$M_r = 24.3 + (35.5 \times 2) + [6 \times ((2 \times 1) + 16)]$
$= 203.3 \text{ g mol}^{-1}$

molar mass = ... 203.3 ..... g mol^{-1}

[1]

(ii) What would the student see after adding the aqueous silver nitrate, $AgNO_3(aq)$?

White precipitate

[1]

(iii) Write an ionic equation for this reaction. Include state symbols.

$Ag^+(aq) + Cl^-(aq) \rightarrow AgCl(s)$

[2]

(iv) Using aqueous silver nitrate, it is sometimes difficult to distinguish between chloride, bromide and iodide ions.

How can aqueous ammonia be used to distinguish between these three ions?

AgCl dissolves in $NH_3(aq)$

AgBr " " conc. $NH_3(aq)$

AgI insoluble in $NH_3(aq)$

[3]

[Total 7 marks]

18. In 2000, the mass of CO_2 emitted in the UK was equivalent to 1 kg per person in every hour.

$M_r(CO_2) = 12 + (16 \times 2) = 44$

(i) Calculate the volume of 1 kg of carbon dioxide. Assume that 1 mole of CO_2 occupies 24 dm^3.

$n(CO_2) = \frac{1000 \text{ g}}{44}$
$= 22.73 \text{ mol} \times 24$
$= 545.5 \text{ dm}^3$

volume = ... 545 ..... dm^3

[2]

(ii) The UK has set a target to cut CO_2 emissions by 60% of the 2000 value by 2050. Calculate the reduction needed in the volume of CO_2 emissions each hour per person if the target is to be met.

Reduction = $545 \times \frac{60}{100}$

= 327

answer: 327 dm^3

[1]

[Total 3 marks]

$Ca + 2HCl \longrightarrow CaCl_2 + H_2$

19. To prepare the aqueous calcium chloride, the student added the exact amount of calcium so that all the hydrochloric acid had reacted. She used 50 cm^3 of 2.0 mol dm^{-3} HC*l*.

(i) How many moles of HC*l* had she used? $n(HCl) = \frac{50}{1000} \times 2.0 = 0.1$ mol

[1]

(ii) Calculate the mass of calcium that she used.

$n(Ca) = \frac{n(HCl)}{2} = 0.1$ mol

Mass = $\frac{0.1}{2} \times 40.1 = 2.005$ g

[2]

(iii) The student added some more calcium and she was surprised that a reaction still took place.

- Explain this observation.
- Write a balanced equation for this reaction.

Ca has reacted with water

$Ca + 2H_2O \longrightarrow Ca(OH)_2 + H_2$

[3]

[Total 6 marks]

20. A student reacted 1.44 g of titanium with chlorine to form 5.70 g of a chloride **X**.

(i) How many moles of Ti atoms were reacted? $n(Ti) = \frac{1.44}{47.9} = 0.03$ mol

[1]

(ii) How many moles of C*l* atoms were reacted?

Mass of Cl = 5.70 − 1.44 = 4.26

$n(Cl) = \frac{4.26}{35.5} = 0.12$ mol

[2]

(iii) Determine the empirical formula of **X**.

Ti : Cl

0.03 0.12

$\frac{0.03}{0.03}$ $\frac{0.12}{0.03}$ 1 : 4

$TiCl_4$

[1]

(iv) Construct a balanced equation for the reaction between titanium and chlorine.

$Ti + 2Cl_2 \longrightarrow TiCl_4$

[1]

[Total 5 marks]

21. A student had a stomach-ache and needed to take something to neutralise excess stomach acid. He decided to take some Milk of Magnesia, which is an aqueous suspension of magnesium hydroxide, $Mg(OH)_2$.

(a) The main acid in the stomach is hydrochloric acid, HC*l*(aq), and the unbalanced equation for the reaction that takes place with Milk of Magnesia is shown below.

.....$Mg(OH)_2$(s) + ..2..HC*l*(aq) → .....$MgCl_2$(aq) + ..2..H_2O(l)

Balance the equation by adding numbers where necessary in the unbalanced equation above.

[1]

(b) The student's stomach contained 500 cm^3 of stomach fluid with an acid concentration of 0.108 mol dm^{-3}. The student swallowed some Milk of Magnesia containing 2.42 g $Mg(OH)_2$. He wondered whether this dose was sufficient to neutralise the stomach acid.

Assume that all the acid in the stomach fluid was 0.108 mol dm^{-3} hydrochloric acid.

(i) How many moles of HC*l* were in the 500 cm^3 of stomach fluid?

$n(HCl) = \frac{500}{1000} \times 0.108 = 0.054$ mol [1]

(ii) Calculate the mass of $Mg(OH)_2$ necessary to neutralise this stomach fluid.

$n(Mg(OH)_2) = \frac{0.108}{2} = 0.027$ Mass = $0.027 \times 58.3 = 1.57$g [3]

(iii) Determine whether the student swallowed too much, too little, or just the right amount of Milk of Magnesia to neutralise the stomach acid.

Too much as dose = 2.42g

.....

[1]

[Total 6 marks]

22. Bromine forms three compounds with phosphorus. The compounds have the molecular formulae PBr_3, PBr_5 and P_2Br_4.

(i) Explain what is meant by the term *molecular formula.*

Number and type of atoms making up a molecule

.....

[1]

(ii) PBr_3 can be prepared by heating bromine with phosphorus, P_4.

Write a balanced equation for this reaction.

$P_4 + 6\,Br_2 \longrightarrow 4\,PBr_3$

[1]

(iii) Compound **A** is one of the three bromides of phosphorus above. It has the following percentage composition by mass: P, 16.2%; Br, 83.8%.

Use this percentage composition to calculate the empirical formula and to determine the identity of compound **A**.

P	Br
$\frac{16.2}{31}$	$\frac{83.8}{79.9}$
0.52	1.05
$\frac{0.52}{0.52}$	$\frac{1.05}{0.52}$
1	2

empirical formula PBr_2

identity of compound **A** P_2Br_4

[3]

[Total 5 marks]

23. Nickel makes up 25% of the total mass of a fifty pence coin. A fifty pence coin has mass of 8.0 g.

(i) Calculate how many **moles** of nickel atoms are in a fifty pence coin.

Mass Ni = $8.0 \times \frac{25}{100} = 2.0\,g$

$n(Ni) = \frac{2.0}{58.7} = 0.034$ moles

answer 0.034 mol

[2]

(ii) Calculate the **number** of atoms of nickel in a fifty pence coin.

$L = 6.02 \times 10^{23}\ mol^{-1}$

N^o atoms $= 0.034 \times 6.02 \times 10^{23}$

$= 2.05 \times 10^{22}$

answer 2.1×10^{22} atoms

[1]

[Total 3 marks]