Chemistry: Chapter 12 Reacting masses
Combined Science (Chemistry Part): Chapter 12 Reacting masses

Section 12.1
|!|ELA031212001O|!|
Calculate the masses of the following:
(a) 1.5 moles of methane CH_4

(b) 0.22 mole of glucose $C_6H_{12}O_6$

(c) 12 moles of sodium chloride NaCl

(d) 2.5 moles of zinc ions Zn^{2+}

(e) 3 moles of sodium carbonate-10-water $Na_2CO_3 \cdot 10H_2O$

[5M]

##
(a) Molar mass of CH_4 = 12.0 + 1.0 × 4 g mol^{-1} = 16.0 g mol^{-1}
Mass of 1.5 moles CH_4 = 16.0 × 1.5 g = 24.0 g [1]
(b) Molar mass of $C_6H_{12}O_6$ = 12.0 × 6 + 1.0 × 12 + 16.0 × 6 g mol^{-1} = 180.0 g mol^{-1}
Mass of 0.22 mole $C_6H_{12}O_6$ = 180.0 × 0.22 g = 39.6 g [1]
(c) Molar mass of NaCl = 23.0 + 35.5 g mol^{-1} = 58.5 g mol^{-1}
Mass of 12 moles NaCl = 58.5 × 12 g = 702.0 g [1]
(d) Molar mass of Zn^{2+} = 65.4 g mol^{-1}
Mass of 2.5 moles Zn^{2+} = 65.4 × 2.5 g = 163.5 g [1]
(e) Molar mass of $Na_2CO_3 \cdot 10H_2O$
= 23.0 × 2 + 12.0 + 16.0 × 3 + 10 × (1.0 × 2 + 16.0) g mol^{-1} = 286.0 g mol^{-1}
Mass of 3 moles of $Na_2CO_3 \cdot 10H_2O$ = 286.0 × 3 g = 858.0 g [1]
##

|!|ELA031212002O|!|
Given 9.6 g of ammonium carbonate $(NH_4)_2CO_3$, find
(a) the number of moles of the compound.

(b) the number of moles of ammonium ions.

(c) the number of moles of carbonate ions.

(d) the number of moles of hydrogen atoms.

(e) the number of hydrogen atoms.

[5M]

##
(a) Molar mass of $(NH_4)_2CO_3$ = (14.0 + 1.0 × 4) × 2 + 12.0 + 16.0 × 3 g mol^{-1}= 96.0 g mol^{-1}

Number of moles of $(NH_4)_2CO_3$ molecules = $\frac{9.60}{96.0}$ mol = 0.100 mol [1]

(b) Since one formula unit of $(NH_4)_2CO_3$ contains two NH_4^+ ions,
number of moles of NH_4^+ ions = 0.1 × 2 mol = 0.2 mol [1]

(c) Since one formula unit of $(NH_4)_2CO_3$ contains one CO_3^{2-} ion,
number of moles of CO_3^{2-} ions = 0.1 mol [1]

(d) Since one formula unit of $(NH_4)_2CO_3$ contains eight H atoms,
number of moles of H atoms = 0.1 × 8 mol= 0.8 mol. [1]

(e) Number of H atoms = 0.8 × 6.02 × 10^{23} = 5 × 10^{23} [1]
##

|!|ELB031212003O|!|

(a) What is the molar mass of water? Is it easy to measure out one mole of water?

(b) What is the mass of one water molecule?

(c) If chemists had chosen 1 × 10^{18} as the number of particles in one 'mole', what would the 'molar mass' of water molecules then be? In that case, is it easy to measure out one 'mole' of water?

[5M]

##
(a) Molar mass of H_2O = 1.0 × 2 + 16.0 mol = 18.0 g mol^{-1} [1]
Yes. [1]

(b) Mass of 1 H_2O molecule = $\frac{18.0}{6.02\times10^{23}}$ g = 2.99 × 10^{-23} g [1]

(c) If 1 'mole' has 1 × 10^{18} molecules, then, the molar mass of water

$= 18.0 \times \dfrac{1\times10^{18}}{6.02\times10^{23}}$ g mol^{-1} = 2.99 × 10^{-5} gmol^{-1} [1]

No. [1]

##

|!|ELA031212004O|!|

Calculate the following quantities:

(a) The number of moles of molecules in 3.4 g of ammonia.

(b) The number of moles of atoms in 3.01 × 10^{24} atoms of gold.

(c) The total number of moles of atoms in 60 g of glucose ($C_6H_{12}O_6$).

(d) The number of moles of
(i) compound,

(ii) ions in 2.84 g of sodium sulphate (Na_2SO_4).

[6M]

##

(a) Molar mass of NH_3 = 14.0 + 1.0 × 3 g mol^{-1} = 17.0 g mol^{-1}

Number of moles of molecules in 3.4 g NH_3 = $\dfrac{3.4}{17.0}$ mol = 0.20 mol [1]

(b) Number of moles of atoms in 3.01 × 10^{24} gold atoms

$= \dfrac{3.01\times10^{24}}{6.02\times10^{23}}$ mol = 5.00 mol [1]

(c) Molar mass of $C_6H_{12}O_6$ = 12.0 × 6 + 1.0 × 12 + 16.0 × 6 g mol^{-1}= 180.0 g mol^{-1}

Number of moles of molecules in 60 g $C_6H_{12}O_6$ = $\dfrac{60}{180.0}$ mol = 0.33 mol [1]

Since 1 formula unit of $C_6H_{12}O_6$ molecule contains 6 + 12 + 6 = 24 atoms, total number of moles of atoms in 0.33 mole $C_6H_{12}O_6$ molecule = 24 × 0.33 mol = 7.9 mol [1]

(d) (i) Molar mass of Na_2SO_4 = 23.0 × 2 + 32.1 + 16.0 × 4 g mol^{-1}= 142.1 g mol^{-1}
Number of moles of compounds in 2.84 g of Na_2SO_4

$= \dfrac{2.84}{142.1}$ mol = 0.0200 mol [1]

(ii) Since one formula unit of Na_2SO_4 has three ions, number of moles of ions in

0.02 mole Na_2SO_4 = 3 × 0.0200 mol = 0.0600 mol [1]
##

|!|ELB031212005O|!|
(a) If 1 g of $^{16}_{8}O$ contains y atoms, how many atoms does 7 g of $^{14}_{7}N$ contain? (Express your answer in terms of y.)

(b) If 1 mole of XO_2 contains the same number of atoms as 60 g of XO_3, what is the molar mass of XO_2?

[8M]
##
(a) Number of moles of atoms in 1 g $^{16}_{8}O$ = $\frac{1}{16.0}$ mol [1]

$\frac{1}{16.0}$ × Avogadro's constant = y
Avogadro's constant = $16y$ [1]
Number of moles of atoms in 7 g $^{14}_{7}N$ = $\frac{7}{14.0}$ mol = 0.5 mol [1]
Number of atoms in 7 g $^{14}_{7}N$ = 0.5 × $16y$ = $8y$ [1]
(b) Let z be the relative atomic mass of element X, and L be the Avogadro's constant.
Number of atoms in 60 g XO_3
= $\frac{60}{z+48}$ × L × 4 [1]
Number of atoms in 1 mol XO_2
= 1 × L × 3 [1]
$1 \times L \times 3 = \frac{60}{z+48} \times L \times 4$
z = 32 [1]
Hence, the molar mass of XO_2 is 64 g mol^{-1}. [1]
##

Section 12.2
|!|ELA031212006O|!|
The bromide of a metal X has the formula XBr_2 and contains 50% by mass of X.
Find the relative atomic mass of X.

[3M]

##

Let x be the relative atomic mass of *X*.

Fraction by mass of *X* in XBr_2 = $\frac{\text{relative atomic mass of } X \times 1}{\text{formula mass of } XBr_2}$ [1]

$$\frac{50}{100} = \frac{x}{x + 79.9 \times 2}$$ [1]

$x = 159.8$ [1]

##

Section 12.3

|!|ELA031212007O|!|

(a) Find the percentage by mass of sodium, sulphur, oxygen and hydrogen in sodium sulphate-10-water, $Na_2SO_4 \cdot 10H_2O$.

(b) Find the mass of water of crystallization in 64.4 g of sodium sulphate-10-water.

(c) Find the total mass of metal(s) in

(i) 20 g of $FeSO_4 \cdot 7H_2O$.

(ii) 100 g of $KMnO_4$.

(iii) 50 g of $K_2Cr_2O_7$.

(d) 2.8 g of a metal *M* combines with 1.2 g of oxygen to form an oxide with formula M_2O_3. What is the relative atomic mass of *M*?

[13M]

##

(a) Formula mass of $Na_2SO_4 \cdot 10H_2O$

$= 23.0 \times 2 + 32.1 + 16.0 \times 4 + 10 \times (1.0 \times 2 + 16.0)$ g mol^{-1} = 322.1 g mol^{-1}

% by mass of Na = $\frac{23.0 \times 2}{322.1} \times 100\% = 14.3\%$ [1]

% by mass of S = $\frac{32.1}{322.1} \times 100\% = 9.97\%$ [1]

% by mass of O = $\frac{16.0 \times 14}{322.1} \times 100\% = 69.5\%$ [1]

% by mass of H = $\frac{1.0 \times 20}{322.1} \times 100\% = 6.2\%$ [1]

(b) Mass of water of crystallization

= mass of $Na_2SO_4 \cdot 10H_2O$ × fraction by mass of water in $Na_2SO_4 \cdot 10H_2O$

$= 64.4 \text{ g} \times \frac{18.0 \times 10}{322.1}$

= 36.0 g [1]

(c) (i) Molar mass of $FeSO_4 \cdot 7\ H_2O$

$= 55.8 + 32.1 + 16.0 \times 4 + 7 \times (1.0 \times 2 + 16.0)$ g mol^{-1} = 277.9 g mol^{-1}

Total mass of metal in 20 g $FeSO_4 \cdot 7\ H_2O$

$= 20 \text{ g} \times \frac{55.8}{277.9} = 4.02$ g [1]

(ii) Molar mass of $KMnO_4 = 39.1 + 54.9 + 16.0 \times 4$ g mol^{-1} = 158.0 g mol^{-1}

Total mass of metal in 100 g $KMnO_4 = 100 \text{ g} \times \frac{39.1 + 54.9}{158.0} = 59.5$ g [1]

(iii) Molar mass of $K_2Cr_2O_7$

$= 39.1 \times 2 + 52.0 \times 2 + 16.0 \times 7$ g mol^{-1} = 294.2 g mol^{-1}

Total mass of metal in 50 g $K_2Cr_2O_7$

$= 50 \text{ g} \times \frac{39.1 \times 2 + 52.0 \times 2}{294.2} = 31.0$ g [1]

(d) $4M(s) + 3O_2(g) \rightarrow 2M_2O_3(s)$ [0.5]

Number of moles of O_2 used = $\frac{1.2}{32.0}$ mol = 0.0375 mol [1]

From the equation, mole ratio of $M : O_2 = 4 : 3$ [0.5]

Number of moles of $M = \frac{0.0375}{3} \times 4$ mol = 0.0500 mol [1]

Let the relative atomic mass of M be y, $\frac{2.8}{y} = 0.05$

$\therefore y = 56$ [1]

Therefore, the relative atomic mass of M is 56. [1]

##

<u>Section 12.4</u>

|!|ELA031212008O|!|

5.8 g of a dry gaseous compound X (containing carbon and hydrogen only) were completely burnt in excess dry oxygen. The products were passed through a suitable drying apparatus containing anhydrous calcium chloride, and it was found that 9.0 g of water had been formed.

(a) Calculate the mass of hydrogen in 9.0 g of water.

(b) Calculate the mass of carbon in 5.8 g of *X*.

(c) Find the empirical formula of *X*.

[5M]

##

(a) Formula mass of H_2O = 1.0 × 2 + 16.0 = 18.0 g

Mass of H in 9.0 g of water = $9.0 \times \frac{1.0 \times 2}{18.0}$ g = 1.0 g [1]

(b) Mass of C in 5.8 g of *X* = (5.8 − 1.0) g = 4.8 g [1]

(c) Since all the C in CO_2 and H in H_2O came from the same compound,

	C	H
Masses (in g)	4.8	1.0
Number of moles (mol)	$\frac{4.8}{12.0}$ = 0.4	$\frac{1.0}{1.0}$ = 1.0
Relative number of moles	$\frac{0.4}{0.4}$ = 1	$\frac{1.0}{0.4}$ = 2.5
Simplest whole number ratio	1 × 2 = 2	2.5 × 2 = 5

[2]

∴ The empirical formula of compound *X* is C_2H_5. [1]

##

|!|ELA031212009O|!|

A compound *X* containing only carbon, hydrogen and oxygen gave the following results on analysis: 3.72 g of the substance gave 5.28 g of carbon dioxide and 3.24 g of water on complete combustion.

(a) Find the mass of carbon, hydrogen and oxygen in the sample of *X*.

(b) Find the empirical formula of compound X.

[7M]

##

(a) Since all the C in CO_2 and H in H_2O came from the same compound X,

mass of C in the compound X = 5.28 g × $\frac{12.0}{12.0+16.0\times 2}$ = 1.44 g [1]

mass of H in the compound X = 3.24 g × $\frac{2.0}{1.0\times 2+16.0}$ = 0.36 g [1]

mass of O in the compound X = (3.79 − 1.44 − 0.36) g = 1.99 g [1]

(b)

	C	H	O
Masses (in g)	1.44	0.36	1.92
Number of moles (mol)	$\frac{1.44}{12.0}$ = 0.12	$\frac{0.36}{1.0}$ = 0.36	$\frac{1.92}{16.0}$ = 0.12
Relative number of moles	$\frac{0.12}{0.12}$ = 1	$\frac{0.36}{0.12}$ = 3	$\frac{0.12}{0.12}$ = 1

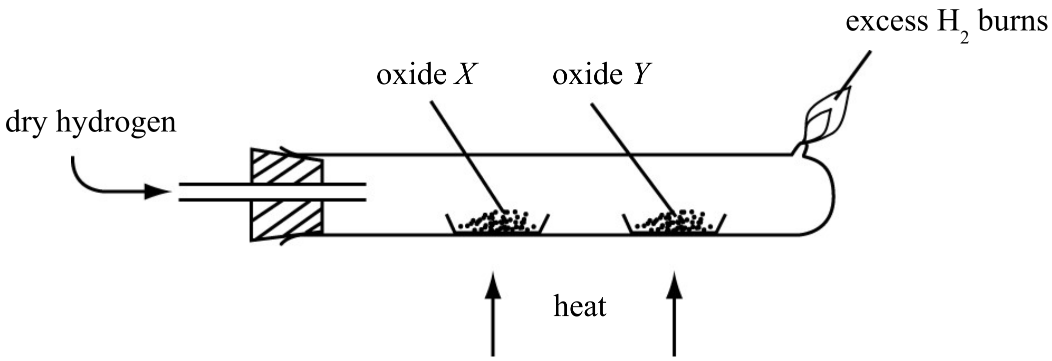
[3]

The empirical formula of compound X is CH_3O. [1]

##

|!|ELB0312120010O|!|

Dry hydrogen is passed over two heated oxides X and Y of the same metal M as shown below.



It is found that 5.4 g of X and 6.0 g of Y are both reduced to 4.2 g of M. If the empirical formula of Y is M_2O_3, work out the empirical formula of X.

[7M]

##

Let z be the relative atomic mass of M,

$2M_2O_3(s) \rightarrow 4M(s) + 3O_2(g)$

6.0 g 4.2 g 1.8 g [0.5]

Number of moles of M_2O_3 used = $\frac{6.0}{2z+48}$ mol [1]

Number of moles of M = $\frac{4.2}{z}$ mol [1]

From the equation, the mole ratio of $M_2O_3 : M = 2 : 4$ [0.5]

So, $\frac{6.0}{2z+48} = \frac{4.2}{z} \times \frac{1}{2}$

$6.0z = 4.2z + 100.8$

$1.8z = 100.8$

$z = 56$ [1]

For the reduction of oxide X,

	M	O
Masses (in g)	4.2	5.4 – 4.2 = 1.2
Number of moles	$\frac{4.2}{56.0} = 0.075$	$\frac{1.2}{16.0} = 0.075$
Relative number of moles	$\frac{0.075}{0.075} = 1$	$\frac{0.075}{0.075} = 1$

[2]

The empirical formula for oxide X is MO. [1]

##

<u>Sections 12.5−12.6</u>

|!|ELB031212011O|!|

Copper can be extracted from copper ore containing copper(II) sulphide, by carbon reduction.

(a) Suggest a solid reducing agent used in the extraction.

____

(b) Write chemical equation(s) to illustrate the reaction(s) in the extraction process.

____

(c) Explain why lead extraction requires a higher temperature than copper extraction 6does.

(d) Suggest a method to extract copper except carbon reduction.

(e) How many grams of reducing agent suggested in part (a) is required to extract all copper from 8.604 g of copper(II) sulphide?

[11M]

##
(a) Carbon / coke [1]
(b) $2CuS(s) + 3O_2(g) \rightarrow 2CuO(s) + 2SO_2(g)$ [1]
$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$ [1]
(c) Lead is more reactive than copper, [1] thus lead compounds are more stable than copper compounds. [1] As a result, more energy, and hence higher temperature, is needed to convert lead compounds into lead. [1]
(d) Add zinc powder into copper(II) sulphate solution. [1]
Or any correct displacement reactions.
(e) Consider the following equations:
$2CuS(s) + 3O_2(g) \rightarrow 2CuO(s) + 2SO_2(g)$ …………..(I)
$2CuO(s) + C(s) \rightarrow 2Cu(s) + CO_2(g)$ ……………….(II)
Number of moles of CuS(s) present = $\frac{8.604}{63.5+32.1}$ mol = 0.09 mol [1]
From equation (I), mole ratio of CuS to CuO = 1:1
$\therefore$ number of moles of CuO formed = 0.09 mol [1]
From equation (II), mole ratio of CuO to C = 2:1
⇦ number of moles of C needed = 0.045 mol [1]
Mass of C needed = 0.045 × 12.0 g = 0.54 g [1]
##

|!|ELB0312120120|!|
In an experiment, a sample of 24.08 g iron burns in oxygen to form an oxide.
(a) What is the formula of the oxide if the oxide contains 70% of iron by mass?

(b) Find the formula mass of the oxide.

(c) What is the mass of oxygen required for a complete combustion of the iron sample? Write down all steps in your calculations.

(d) The oxide is then reduced back to iron by carbon. Calculate the amount of carbon needed to obtain all iron in the product formed from the reaction stated in part (c).

[12M]

##

(a) Let there be 100 g of the iron oxide.

Mole ratio of Fe : O = $\frac{70}{55.8}$: $\frac{100-70}{16.0}$ [1]

= 1.25 : 1.875

= 2 : 3 [1]

The formula of the oxide is Fe_2O_3. [1]

(b) The formula mass of Fe_2O_3

= 2 × 55.8 + 3 × 16

= 159.6 [1]

(c) The equation for the reaction is

	$4Fe(s)$ +	$3O_2(g)$ →	$2Fe_2O_3(s)$
Mole ratio	4	3	
Number of moles (mol)	$\frac{24.08}{55.8}$ = 0.43 [1]	$0.43 \times \frac{3}{4}$ = 0.33 [1]	

[3]

The mass of oxygen required for complete combustion

= 0.33 × (16.0 × 2)

= 10.56 g [1]

(d) Refer to the equation in part (c), the amount of Fe_2O_3 produced is $\frac{0.43}{2}$ mol = 0.22 mol. [1]

The equation for the reduction is

	$2Fe_2O_3(s)$ +	$3C(s)$ →	$4Fe(s)$ +	$3CO_2(g)$
Mole ratio	2	3		
Number of mole (mol)	0.22	$0.22 \times \frac{3}{2}$ = 0.33		

[1]

The mass of carbon required for complete reduction
= 0.33 × 12.0 g
= 3.96 g [1]

##

|!|ELB031212013O|!|

Silver is formed if zinc powder or magnesium powder is added to silver nitrate solution.

(a) Adding which powder, zinc or magnesium, would the process of silver formation be faster?

(b) Explain why magnesium powder helps obtaining silver faster than magnesium ribbon does.

(c) How many grams of magnesium powder should be added into silver nitrate solution to obtain 5.0 g of silver?

(d) If 0.8 g of zinc reacts with excess silver nitrate solution to form silver. What mass of silver can be obtained?

[11M]

##

(a) Magnesium powder [1]. As magnesium is more reactive than zinc [1], magnesium undergoes displacement reaction faster.

(b) Magnesium powder provides greater surface area for the reaction to occur. [1]

(c) The ionic equation for the reaction is

	$Mg(s)$ +	$2Ag^+(aq)$ →	$Mg^{2+}(aq)$ +	$2Ag(s)$
Mole ratio	1			2
Number of moles (mol)	$\frac{0.047}{2}$			$\frac{5.0}{108}$
	= 0.024			= 0.047

[3]

Mass of magnesium required to obtain 5.0 g of silver
= 0.024 × 24.3 g
= 0.58 g [1]

(d) The equation for the reaction is

	$Zn(s)$ +	$2AgNO_3(aq)$ →	$Zn(NO_3)_2(aq)$ +	$2Ag(s)$ [1]
Mole ratio	1			2
Number of moles (mol)	$\frac{0.8}{65.4}$			0.01 × 2
	= 0.01 [1]			= 0.02 [1]

The mass of silver obtained
= 0.02 × 108 g
= 2.6 g [1]

##

|!|ELB031212014O|!|

X, *Y* and *Z* are three different metal solids, having 2+ charge when in compounds. The table below shows the results of two experiments about the metals.

Experiment	**Metal *X***	**Metal *Y***	**Metal *Z***
Adding metal to dilute hydrochloric acid	Gas evolved	No reaction	Gas evolved
Adding metal to cold water	No reaction	No reaction	Gas evolved

(a) Arrange the three metals in ascending order of reactivity.

(b) Which metal forms the most stable oxide? Explain your answer.

(c) Suggest and explain what metal *Y* would be.

(d) When 0.42 g of molten oxide of metal *Y* undergoes electrolysis, it decomposes completely to give 0.12 g of oxygen. Calculate the formula mass of the oxide and the relative atomic mass of *Y*.

[15M]

##
(a) $Y < X < Z$ [2]
(b) Metal *Z*. [1] Metal *Z* is the most reactive and hence its oxide is the most stable. [1]
(c) Copper. [1] As *Y* does not react with both cold water and dilute hydrochloric acid, it must be lower than lead in the reactivity series. [1] The metals lower than lead are copper, mercury, silver, platinum and gold. [1] Mercury is liquid [1], the charge of silver in compound is usually 1+ [1], and gold usually does not form any oxide at all [1]. Copper is a solid and its charge in compound can be 2+ [1]. Therefore, it is reasonable to believe that *Y* is copper.
(d) Number. of mole of oxygen atoms in *Y*O = $\frac{0.12}{16.0}$ mol = 0.0075 mol [1]
Number of mole of *Y*O = Number of mole of oxygen atoms = 0.0075 mol [1]
Formula mass of *Y*O = $\frac{0.42}{0.0075}$ = 56 [1]
Relative atomic mass of *Y* = 56 – 16 = 40 [1]
##

|!|ELA0312120150|!|
A mixture of zinc sulphide and zinc oxide is heated. After strong heating for 30 minutes, 0.52 g of a gaseous product *A* is formed. The residue is then heated strongly with excess carbon. 0.70 g of another gaseous product *B* is formed after the reaction is complete.
(a) Identify the two gaseous products in the reactions involved.

(b) Write chemical equations for the reactions as described above.

(c) Calculate the percentage by mass of zinc oxide in the mixture.

[13M]

##

(a) *A* is sulphur dioxide. *B* is carbon dioxide. [2]

(b) $2ZnS(s) + 3O_2(g) \rightarrow 2ZnO(s) + 2SO_2(g)$ [1]

$2ZnO(s) + C(s) \rightarrow 2Zn(s) + CO_2(g)$ [1]

(c) The first reaction is

	$2ZnS(s)$ +	$3O_2(g)$ →	$2ZnO(s)$ +	$2SO_2(g)$
Mole ratio	1		1	1
Number of moles (mol)	0.008		0.008	$\frac{0.52}{32.1+16.0\times 2}$ =0.008
Mass (g)	0.008 × (65.4 + 32.1) = 0.8 g		0.008 × (65.4 + 16.0) = 0.7 g	

[4]

The second reaction is

	$2ZnO(s)$ +	$C(s)$ →	$2Zn(s)$ +	$CO_2(g)$
Mole ratio	2			1
Number of moles	0.02 × 2 = 0.04			$\frac{0.7}{12.0+16.0\times 2}$ =0.02
Mass	0.04 × (65.4 + 16.0) = 3.3 g			

[2]

Total mass of ZnO(s) reacted is 3.3 g. This includes the ZnO(s) initially present in the mixture and the product from ZnS(s). Therefore, mass of ZnO(s) initially present in the mixture = 3.3 – 0.7 = 2.6 g [1]

The total mass of the mixture = mass of ZnS(s) + mass of ZnO(s)

= 0.78 + 2.6

= 3.4 g [1]

The percentage by mass of ZnO(s) in the mixture
$= \frac{2.6}{3.4} \times 100\%$
= 76.5 % [1]
##

|!|ELA031212016O|!|
Calculate the mass of potassium superoxide formed when 3.91 g of potassium is burnt with 2.56 g of oxygen.

[13M]
##
The chemical equation of this reaction is: $K(s) + O_2(g) \rightarrow KO_2(s)$ [1]
Molar mass of K = 39.1 g mol^{-1}
Number of moles of K = $\frac{3.91}{39.1}$ mol [1]
= 0.1 mol [1]
Molar mass of O_2 = 32.0 g mol^{-1}
Number of moles of O_2 = $\frac{2.56}{32.0}$ mol [1]
= 0.08 mol [1]
From the equation, mole ratio of K : O_2 = 1 : 1 [1]
Hence, 0.08 mol of O_2 would react with 0.08 mol of K. [1]
Since 0.1 mol of K is used, K is in excess. [1]
Hence, number of moles of KO_2 formed = 0.08 mol [1]
Molar mass of KO_2 = (39.1 + 16.0 × 2) [1]
= 71.1 g mol^{-1} [1]
Mass of KO_2 formed = (0.08 × 71.1) g [1]
= 5.7 g [1]
##

|!|ELA031212017O|!|
Calculate the mass of copper produced by the complete reaction of 60.0 g of copper(II) sulphate in the following reaction.
$CuSO_4(aq) + Fe(s) \rightarrow FeSO_4(aq) + Cu(s)$

[8M]

##
Mass of $CuSO_4$ = 60.0 g
Molar mass of $CuSO_4$ = (63.5 + 16.0 × 4 + 32.1) g mol^{-1} [1] = 159.6 g mol^{-1} [1]
Number of moles of $CuSO_4$ = $\frac{60.0}{159.6}$ mol [1] = 0.376 mol [1]
From the chemical equation, mole ratio of $CuSO_4$: Cu = 1 : 1 [1]
Hence, number of moles of Cu produced = 0.376 mol [1]
Mass of copper produced = (0.376 × 63.5) g [1] = 23.9 g [1]
##

|!|ELA031212018O|!|
A student added 46.0 g of sodium to 48.0 g of water and observed what happened.
(a) Write a chemical equation of the reaction involved.

(b) Determine which reactant is in excess.

(c) Calculate the mass of hydrogen that could be obtained.

(d) State what would happen if the resultant solution is tested with red litmus paper.

(e) State how sodium could be stored in school laboratories.

[12M]

##
(a) $2Na(s) + 2H_2O(l) \rightarrow 2NaOH(aq) + H_2(g)$ [1]
(b) Number of moles of Na = $\frac{46.0}{23.0}$ mol [1] = 2 mol [1]

Number of moles of H_2O = $\frac{48.0}{2.0+16.0}$ mol [1] = 2.7 mol [1]

From the equation, 2 moles of sodium react with 2 moles of water. [1] Hence, 0.7 moles of water is in excess. [1]

(c) As the mole ratio of Na : H_2 = 2 : 1 and the limiting reactant is Na, 1 mole of hydrogen would be produced. [1]
Hence, the maximum mass of hydrogen produced = (1.0 × 2) g [1] = 2.0 g [1]

(d) Red litmus paper turns blue. [1]

(e) Sodium could be stored under paraffin oil. [1]

##

|!|ELA031212019O|!|

A student was given 32.7 g of zinc granules and all the granules were added to dilute hydrochloric acid.

(a) Write a chemical equation for the reaction involved.

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(b) Determine the amount of dilute hydrochloric acid needed for all zinc granules to react completely.

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(c) Calculate the mass of zinc chloride that could be formed.

____

(d) Was this reaction exothermic or endothermic?

____

[14M]

##

(a) $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$ [1]

(b) Number of moles of Zn used= $\frac{32.7}{65.4}$ mol [1] = 0.5 mol [1]
Molar mass of HCl = (1.0 + 35.5) g mol^{-1} [1] = 36.5 g mol^{-1} [1]
From the equation, mole ratio of Zn : HCl = 1 : 2. [1]
Hence, 0.5 mol of zinc require 0.5 × 2 = 1 mol of HCl for reaction. [1]
Amount of HCl required = 36.5 g [1]

(c) Molar mass of $ZnCl_2$ = (65.4 + 2 × 35.5) g mol^{-1} [1] = 136.4 g mol^{-1} [1]
From the equation, 0.5 moles of $ZnCl_2$ would be produced. [1]
Mass of $ZnCl_2$ produced = 0.5 × 136.4 g [1] = 68.2 g [1]

(d) It was exothermic. [1]
##