

Suggested Answers on Note (Chapter 10) P.1

Elements	H	C
Relative atomic masses	1	12
One mole weighs	1g	12g
Number of atoms in 1 mole	6×10^{23}	6×10^{23}
Number of atoms in 2 mole	$2 \times 6 \times 10^{23}$	$2 \times 6 \times 10^{23}$

Formula	Term to describe its mass	Molecular or Formula mass
H ₂ O	Molecular / Formula	$(1) \times 2 + (16) \times 1 = 18$
Ca(OH) ₂	Formula	74
H ₂ SO ₄	Molecular / Formula	98
Na ₂ CO ₃ .10H ₂ O	Formula	286

Suggested Answers on Note (Chapter 10) P.3

Substance	Symbol / chemical formula	Relative atomic mass(es)	Formula mass / relative molecular mass	Molar mass (mass of 1 mole of substance)
Magnesium	Mg	Mg = 24.3	$1 \times 24.3 = 24.3$	24.3 g mol ⁻¹
Nitrogen gas	N ₂	N = 14.0	$2 \times 14.0 = 28.0$	28.0 g mol ⁻¹
Sulphur dioxide	SO ₂	O = 16.0 S = 32.1	$2 \times 16.0 = 32.0$ $1 \times 32.1 = 32.1$ Total = 64.1	64.1 g mol ⁻¹
Iron(III) sulphate	Fe ₂ (SO ₄) ₃	O = 16.0 S = 32.1 Fe = 55.8	$12 \times 16.0 = 192.0$ $3 \times 32.1 = 96.3$ $2 \times 55.8 = 111.6$ Total = 399.9	399.9 g mol ⁻¹

Examples:

1. **Molar mass of carbon dioxide = $12 + 16 \times 2 = 44 \text{ g}$**
Mass of carbon dioxide = 7.5×44
= 330 g
2. **Number of mole of Mg = $2.4 / 24 = 0.1$**
Number of Mg atom = $0.1 \times 6 \times 10^{23}$
= 6×10^{22}
3. **Number of Na atoms = $1.5 \times L$**
= $1.5 \times 6 \times 10^{23}$
= 9×10^{23}
4. **Number of moles of zinc ions = $3 \times 10^{24} / 6 \times 10^{23}$**
= 5
There are 5 moles of ions in 3×10^{24} zinc ions.
5. (a) **Number of SO_2 molecules = $0.35 \times 6 \times 10^{23}$**
= 2.1×10^{23}
 (b) **One SO_2 molecule contains 1 S atom and 2 O atoms**
Number of atoms present = $3 \times 2.1 \times 10^{23}$
= 6.3×10^{23}

Suggested Answers on Note (Chapter 10) P.4 – 9

1.

Substance	Chemical formula	Relative atomic masses	Formula mass / relative molecular mass	Molar mass
Hydrogen chloride	HCl	H = 1.0 Cl = 35.5	36.5	36.5 g mol^{-1}
Ethanoic acid	CH_3COOH	H = 1.0 C = 12.0 O = 16.0	60.0	60.0 g mol^{-1}
Aluminium hydroxide	$\text{Al}(\text{OH})_3$	H = 1.0 O = 16.0 Al = 27.0	78.0	78.0 g mol^{-1}
Magnesium carbonate	MgCO_3	C = 12.0 O = 16.0 Mg = 24.3	84.3	84.3 g mol^{-1}

2. Number of mole of oxygen = $20 / (16 \times 2)$
 = 0.625

3. (a) Molar mass of $C_xH_{2x} = 210 / 2.5$
 $= 84 \text{ g}$
- (b) $84 = 12x + 2x$
 $x = 6$
 therefore, the chemical formula = C_6H_{12}
4. (a) Number of moles of Cu = $12.7 / 63.5 = 0.2$
 Number of copper atoms = $0.2 \times 6 \times 10^{23} = 1.2 \times 10^{23}$
- (b) Number of moles of iron = $3 \times 10^{22} / 6 \times 10^{23} = 0.05$
 Mass of iron = $0.05 \times 55.8 = 2.79 \text{ g}$
5. (a) no. of moles of ammonia = 0.25
 no. of ammonia molecules = $0.25 \times 6 \times 10^{23} = 1.5 \times 10^{23}$
- (b) One NH_3 molecule contains 1 N atom and 3 H atoms
 no. of N atoms = 1.5×10^{23}
- (c) no. of H atoms = $1.5 \times 10^{23} \times 3 = 4.5 \times 10^{23}$
- (d) total number of atoms = $1.5 \times 10^{23} + 4.5 \times 10^{23} = 6 \times 10^{23}$
6. (a) no. of mole of $MgCl_2 = 1.2 \times 10^{24} / 6 \times 10^{23} = 2$
 mass of $MgCl_2 = 2 \times (24.3 + 35.5 \times 2) = 190.6 \text{ g}$
- (b) no. of mole of $Mg^{2+} = 2$
 mass of $Mg^{2+} = 2 \times 24.3 = 48.6 \text{ g}$
- (c) no. of mole of $Cl^- = 2 \times 2 = 4$
 mass of $Cl^- = 4 \times 35.5 = 142 \text{ g}$
7. (a) no. of formula units = $4.5 \times 6 \times 10^{23} = 2.7 \times 10^{24}$
- (b) One formula unit of K_2SO_4 contains 2 K^+ ions and 1 SO_4^{2-} ion
 no. of ions = $2.7 \times 10^{24} \times (2 + 1) = 8.1 \times 10^{24}$
8. (a) no. of moles of $CaCl_2 = 3.3 / (40 + 35.5 \times 2) = 0.03$
 no. of formula units of $CaCl_2 = 0.03 \times 6 \times 10^{23} = 1.8 \times 10^{22}$
- (b) no. of chloride ions = $1.8 \times 10^{22} \times 2 = 3.6 \times 10^{22}$
9. No. of H_2O molecules in 1 mole = 6×10^{23}
 Mass of one H_2O molecule = $18 / 6 \times 10^{23} = 3 \times 10^{-23} \text{ g}$
10. (a) Molar mass of Be = 9 g
 1 mole Be atom = 6×10^{23} Be atom = 9 g
 Mass of 1 Be atom = $9 / 6 \times 10^{23} = 1.5 \times 10^{-23} \text{ g}$

$$\begin{aligned} \text{(b) Mass of 1 SO}_2 \text{ molecule} &= (32 + 2 \times 16) / 6 \times 10^{23} \\ &= 1.06 \times 10^{-22} \text{ g} \end{aligned}$$

11.

Substance	Chemical formula	Molar mass of substance (g mol ⁻¹)	Mass of substance present (g)	Number of moles of substance present (mol)	Number of molecules / formula units present
Nitrogen dioxide	NO ₂	46	59.8	1.30	7.8 × 10²³
Lead(II) oxide	PbO	223	44.6	0.2	1.2 × 10²³
Ammonium carbonate	(NH ₄) ₂ CO ₃	96	864	9	5.4 × 10 ²⁴ formula units

12. **Remark:** L is the Avogadro's Number

Substance	Number of mole	Mass (g)	Number of particle
Chlorine molecule (Cl ₂)	2	142	Cl ₂ molecules: 2L Cl atoms: 4L
Phosphorus (P ₄)	0.5	62	P ₄ molecules: L/2 P atoms: 2L
Sulphur (S ₈)	0.5	128	S ₈ molecules: L/2 S atoms: 4L
Sulphur dioxide (SO ₂)	2	128	SO ₂ molecules: 2L S atoms: 2L O atoms: 4L
Sodium hydroxide (NaOH)	2	80	Total no. of ions: 2.4 × 10 ²⁴ = 4L Na ⁺ ions: 2L OH ⁻ ions: 2L
Sulphuric acid (H ₂ SO ₄)	0.5	49	H atoms: L S atoms: L/2 O atoms: 2L H ⁺ ions: L SO ₄ ²⁻ ions: L/2
Sucrose (C ₁₂ H ₂₂ O ₁₁)	0.00833	2.85	C atoms: 6 × 10 ²² = 0.1L H atoms: (0.1/12) × 22L O atoms: (0.1/12) × 11L
Hydrated copper(II) sulphate (CuSO ₄ ·5H ₂ O)	0.5	125	H ₂ O molecules: 5L/2 Cu ²⁺ ions: L/2 SO ₄ ²⁻ ions: L/2
Hydrated sodium carbonate (Na ₂ CO ₃ ·10H ₂ O)	2	572	H ₂ O molecules: 20L Na ⁺ ions: 4L CO ₃ ²⁻ ions: 2L

Suggested Answers on Note (Chapter 10) P.11 – 14

- no. of moles of $\text{CO}_2 = 12 / 24 = 0.5$
mass of $\text{CO}_2 = 0.5 \times (12 + 16 \times 2) = 22.0 \text{ g}$
- no. of moles of $\text{NH}_3 = 6 / 24 = 0.25$
no. of NH_3 molecules = $0.25 \times 6 \times 10^{23} = 1.5 \times 10^{23}$
- no. of moles of $\text{Cl}_2 = 4.56 / 24 = 0.19$
- no. of moles of $\text{O}_2 = 6 / 24 = 0.25$
no. of O_2 molecules = $0.25 \times 6 \times 10^{23} = 1.5 \times 10^{23}$
no. of O atoms = $2 \times 1.5 \times 10^{23} = 3 \times 10^{23}$
- no. of moles of $\text{N}_2 = 4.2 / 28 = 0.15$
volume of $\text{N}_2 = 0.15 \times 24 = 3.6 \text{ dm}^3$

6.

Gas	Mass of 1 mole of gas	Mass of gas present	Number of moles of gas	Volume of gas at r.t.p.
H_2	2.0 g	0.40 g	0.20 mol	4.8 dm^3
CO_2	44.0 g	4.0 g	0.091	2.18 dm^3
SO_2	64.0 g	16.0 g	0.25	6000 cm^3
O_2	32.0 g	39.68 g	1.24 mol	29.76 dm^3

- no. of mole of $\text{CH}_4 = 1 / 24 = 0.0417$
no. of molecules = $0.0417 \times 6 \times 10^{23} = 2.5 \times 10^{22}$

8.

Gas	Molar mass (g mol^{-1})	Mass of gas (g)	No. of moles of gas	No. of gas molecules	Volume of gas at r.t.p. (dm^3)
H_2	2.0	1.0	0.5	3×10^{23}	12.0
O_2	32.0	12.8	0.4	2.4×10^{23}	9.6
CO_2	44.0	22.0	0.5	3×10^{23}	12.0
CH_4	16.0	2.0	0.125	7.5×10^{22}	3.0
SO_3	80.0	20.0	0.25	1.5×10^{23}	6.0
Unknown X	17.0	4.25	0.25	1.5×10^{23}	6.0

9. Number of moles of methane gas
= volume of methane gas / molar volume of gas
= $1.0 \text{ dm}^3 / 24.0 \text{ dm}^3 \text{ mol}^{-1}$
= 0.042 mol

$$\begin{aligned}\text{Number of methane molecules} &= \text{number of moles of methane gas} \times L \\ &= 0.042 \text{ mol} \times 6.02 \times 10^{23} \text{ mol}^{-1} \\ &= 2.5 \times 10^{22}\end{aligned}$$

10. Number of moles of gas
= volume of gas / molar volume of gas
= $600 \text{ cm}^3 / 24\,000 \text{ cm}^3 \text{ mol}^{-1}$
= 0.0250 mol

$$\text{Mass of gas} = (78.98 - 77.53) \text{ g} = 1.45 \text{ g}$$

$$\text{Molar mass of gas} = 1.45 \text{ g} / 0.0250 \text{ mol} = 58.0 \text{ g mol}^{-1}$$

Suggested Answers on Note (Chapter 10) P.16 – 17

1.

Urea

$$\text{CO}(\text{NH}_2)_2 \quad \% \text{ N} = [28 / (12+16+28+4)] \times 100\% = 46.67\%$$

Sodium nitrate

$$\text{NaNO}_3 \quad \% \text{ N} = [14 / (23+14+48)] \times 100\% = 16.47\%$$

Aqueous ammonia

$$\text{NH}_3 \quad \% \text{ N} = [14 / (14+3)] \times 100\% = 82.35\%$$

Ammonium sulphate

$$(\text{NH}_4)_2\text{SO}_4 \quad \% \text{ N} = [28 / (28+8+32+64)] \times 100\% = 21.21\%$$

Ammonium phosphate

$$(\text{NH}_4)_3\text{PO}_4 \quad \% \text{ N} = [42 / (42+12+31+64)] \times 100\% = 28.19\%$$

2. Complete the following table:

Compound	Formula	Relative atomic mass	Formula mass / Relative molecular mass	% by mass of each element in the compound
Sodium chloride	NaCl	Na = 23.1 Cl = 35.5	58.6	% Na = 39.4% % Cl = 60.6%
Water	H₂O	H = 1.0 O = 16.0	18.0	% H = 11.1% % O = 88.9%
Zinc hydroxide	Zn(OH)₂	Zn = 65.4 O = 16.0 H = 1.0	99.4	% Zn = 65.8% % O = 32.2% % H = 2.0%
Sulphuric acid	H₂SO₄	H = 1.0 S = 32.1 O = 16.0	98.1	% H = 2.0% % S = 32.7% % O = 65.2%

3. Formula mass of $\text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 246.4$
 % by mass of water in $\text{MgSO}_4 \cdot 7\text{H}_2\text{O} = 7 \times 18 / 246.4 = 51.1\%$
4. Formula mass of $\text{CaCl}_2 \cdot n\text{H}_2\text{O} = 111.1 + 18n$
 % by mass of water in $\text{CaCl}_2 \cdot n\text{H}_2\text{O} = 18n / (111.1 + 18n) = 49.3\%$
 $\Rightarrow n = 6$
5. Formula mass of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = 106 + 18x$
 no. of mole of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} = 14.3 / (106 + 18x)$
 no. of mole of $\text{H}_2\text{O} = 9 / 18 = 0.5$
 mole ratio of $\text{Na}_2\text{CO}_3 \cdot x\text{H}_2\text{O} : \text{H}_2\text{O} = 1 : x$
 $\therefore [14.3 / (106 + 18x)] / 0.5 = 1 / x$
 $\Rightarrow x = 10$

Suggested Answers on Note (Chapter 10) P.22

2.

Element	Cu	Cl
Mass of element / g	47.4	52.6
No. of mole	$47.4 / 63.5 = 0.75$	$52.6 / 35.5 = 1.48$
Relative no. of mole (Mole ratio)	1	2

The empirical formula = CuCl_2

If the formula mass is 270, then the formula is Cu_2Cl_4

If the formula mass is 540, then the formula is Cu_4Cl_8

Suggested Answers on Note (Chapter 10) P.23 – 27

1.

Element	Mg	C	O
Mass of element / g	28.6	14.3	57.1
No. of mole	1.2	1.2	3.6
Relative no. of mole	1	1	3

The empirical formula = MgCO_3

2.

Element	Pb	O
Mass of element / g	$1.374 - 0.126 = 1.248$	0.126
No. of mole	0.006	0.0079
Relative no. of mole	3	4

The empirical formula = Pb_3O_4

3. % by mass of O in compound X = $100 - 40 - 6.7 = 53.3\%$

Element	C	H	O
Mass of element / g	40	6.6	53.3
No. of mole	3.33	6.7	3.33
Relative no. of mole	1	2	1

\therefore The empirical formula = CH_2O .

Let the molecular formula of compound X be $(\text{CH}_2\text{O})_n$, where n is a positive integer.

Molar mass of $(\text{CH}_2\text{O})_n = 60.0 \text{ g mol}^{-1}$

$(12 + 1 \times 2 + 16) \times n = 60$

$\Rightarrow n = 2$

\therefore The molecular formula of compound X = $(\text{CH}_2\text{O})_2$
= $\text{C}_2\text{H}_4\text{O}_2$

4. In 100 g

Element	C	H
Mass of element / g	85.7	14.3
No. of mole	7.14	14.3
Relative no. of mole	1	2

∴ The empirical formula = CH₂.

Let the molecular formula of compound Y be (CH₂)_n, where n is a positive integer.

$$\text{Molar mass of (CH}_2\text{)}_n = 56.0 \text{ g mol}^{-1}$$

$$(12 + 1 \times 2) \times n = 56$$

$$\Rightarrow n = 4$$

∴ The molecular formula of compound Y = (CH₂)₄
= C₄H₈

5. In 100 g

Element	C	H	O
Mass of element / g	60.0	4.50	35.50
No. of mole	5	4.5	2.22
Relative no. of mole	9	8	4

∴ The empirical formula = C₉H₈O₄.

6. In 100 g

Element	C	H	O
Mass of element / g	26.09	4.35	69.56
No. of mole	2.17	4.35	4.34
Relative no. of mole	1	2	2

The empirical formula = CH₂O₂

Let the molecular formula be (CH₂O)_n

$$\text{Therefore, } n(12 + 1 \times 2 + 16 \times 2) = 46 \Rightarrow n = 1$$

The molecular formula = CH₂O₂

7. (a)

Element	Pb	O
Mass of element / g	62.1	68.5 - 62.1 = 6.4
No. of mole	0.3	0.4
Relative no. of mole	3	4

The empirical formula = Pb₃O₄

(b) 2PbO + PbO₂ = Pb₃O₄

PbO : PbO₂ = 2 : 1

8. In 100 g

Element	H	S	O
Mass of element / g	2.40	39.0	58.60
No. of mole	2.40	1.22	3.66
Relative no. of mole	2	1	3

The empirical formula = H_2SO_3

9. (a) (i) mass of H = $9 \times (2/18) = 1\text{g}$
(ii) mass of C = $5.8 - 1 = 4.8\text{g}$
(iii)

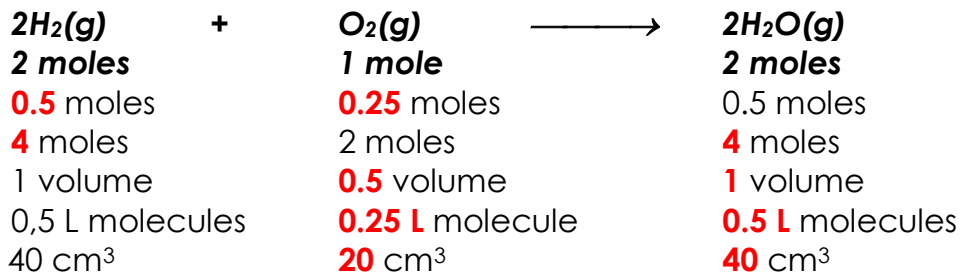
Element	C	H
Mass of element	4.8	1
No. of mole	0.4	1
Relative no. of mole	2	5

The empirical formula = C_2H_5

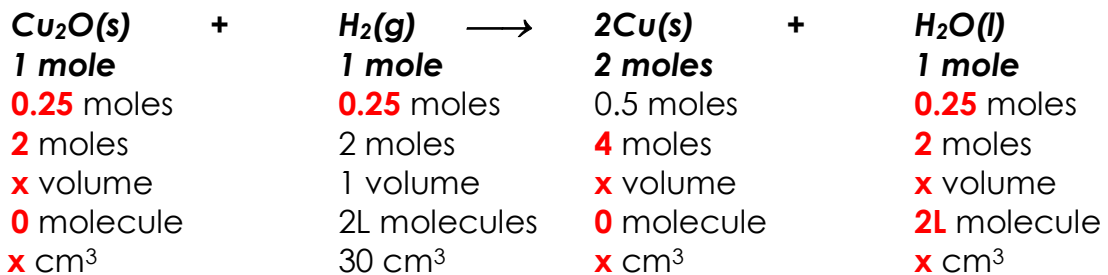
- (b) (i) $1.2/M = 0.5/24 \Rightarrow M = 57.6$ (molecular mass)
(ii) Let the molecular formula = $(\text{C}_2\text{H}_5)_n$
 $n(24+5) = 57.6 \Rightarrow n \sim 2$
Molecular formula = C_4H_{10}

Suggested Answers on Note (Chapter 10) P.29

Example 1:



Example 2:

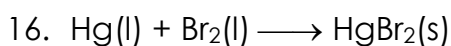


x = very small

Suggested Answers on Note (Chapter 10) P.34 – 42

- $\text{Ca(s)} + 2\text{H}_2\text{O(l)} \longrightarrow \text{Ca(OH)}_2\text{(aq)} + \text{H}_2\text{(g)}$
 Number of mole of calcium provided = $2.1 / 40 = 0.0525$
 According to the equation, 1 mole of Ca reacts with 2 moles of H_2O .
 Number of mole of H_2O needed = $0.0525 \times 2 = 0.105$
 Mass of H_2O needed = $0.105 \times 18 = \mathbf{1.89 \text{ g}}$
- Number of moles of $\text{Fe}_2\text{O}_3 = 63.8 / 159.6 = 0.4$
 Number of moles of Al required = $2 \times 0.4 = 0.8$
 Mass of Al required = $0.8 \times 27 = \mathbf{21.6 \text{ g}}$
- Number of mole of S = 5.5
 Number of mole of $\text{SO}_2 = 5.5$
 Mass of SO_2 formed = $5.5 \times (32 + 16 \times 2) = \mathbf{352 \text{ g}}$
 Volume of SO_2 formed in r.t.p. = $5.5 \times 24 = \mathbf{132 \text{ dm}^3}$
- Number of mole of Pb = $74.6 / 207 = 0.36$
 Number of mole of Pb_3O_4 required = $0.36 / 3 = 0.12$
 Mass of $\text{Pb}_3\text{O}_4 = 0.12 \times 685 = \mathbf{82.3 \text{ g}}$
- $2\text{CuO(s)} + \text{C(s)} \longrightarrow 2\text{Cu(s)} + \text{CO}_2\text{(g)}$
 Number of mole of CuO = $15 / 79.5 = 0.189$
 Number of mole of C provided = $1.5 / 12 = 0.125$
 According to the equation, 2 moles of CuO react with 1 mole of C.
 \therefore 0.125 mole of C needs 0.25 mole of CuO for complete reaction.
 However, there is only 0.189 mole of CuO provided.
 \therefore CuO is the limiting reactant or limiting reagent (or C is in excess).
 Number of moles of Cu formed = 0.189
 Mass of Cu = $0.189 \times 63.5 = \mathbf{12 \text{ g}}$
 Number of moles of $\text{CO}_2 = \frac{1}{2} \times 0.189 = 0.0945$
 Volume of CO_2 formed = $0.0945 \times 24 = \mathbf{2.268 \text{ dm}^3}$
- $\text{Mg(s)} + 2\text{HCl(aq)} \longrightarrow \text{MgCl}_2\text{(aq)} + \text{H}_2\text{(g)}$
 Number of mole of Mg = 0.144
 Number of mole of $\text{MgCl}_2 = 0.144$
 Mass of MgCl_2 formed = $0.144 \times (24.3 + 35.5 \times 2) = \mathbf{13.72 \text{ g}}$
 Number of mole of $\text{H}_2\text{(g)} = 0.144$
 Volume of H_2 formed = $0.144 \times 22.4 = \mathbf{3.23 \text{ dm}^3}$
- $2\text{Al}_2\text{O}_3\text{(l)} \longrightarrow 4\text{Al(l)} + 3\text{O}_2\text{(g)}$
 Number of mole of Al = $5.4 / 27 = 0.2$
 Number of mole of Al_2O_3 reacted = $0.2 / 4 \times 2 = 0.1$
 Mass of Al_2O_3 reacted = $0.1 \times (27 \times 2 + 16 \times 3) = \mathbf{10.2 \text{ g}}$
 Number of mole of O_2 formed = $0.2 / 4 \times 3 = 0.15$
 Volume of O_2 formed = $0.15 \times 24 = \mathbf{3.6 \text{ dm}^3}$

8. Number of mole of sodium carbonate = $5.3 / 106 = 0.05$
 Number of mole of HCl = $2 \times 0.04 = 0.08$
 Sodium carbonate is in excess!!!
 Number of mole of carbon dioxide = $0.08 / 2 = 0.04$
 Volume of carbon dioxide = $0.04 \times 24 = \mathbf{0.96 \text{ dm}^3}$
9. Number of mole of copper(II) oxide = $15.9 / 79.5 = 0.2$
 Number of mole of hydrogen = 0.2
 Mass of copper = $0.2 \times 63.5 = \mathbf{12.7 \text{ g}}$
10. Number of mole of sulphuric acid = $3 \times 0.025 = 0.075$
 Number of mole of hydrated copper(II) sulphate = 0.075
 Molar mass of hydrated copper(II) sulphate = 249.5 g
 Mass of hydrated copper(II) sulphate = $0.075 \times 249.5 = \mathbf{18.1725 \text{ g}}$
11. $\text{NH}_4\text{Cl} + \text{NaOH} \longrightarrow \text{NH}_3 + \text{NaCl} + \text{H}_2\text{O}$
 $2\text{NH}_3 + \text{H}_2\text{SO}_4 \longrightarrow (\text{NH}_4)_2\text{SO}_4$
 Number of mole of ammonium sulphate = $2.64 / 132 = 0.02$
 Number of mole of ammonia = $0.02 \times 2 = 0.04$
 Number of mole of ammonium chloride = 0.04
 Mass of ammonium chloride = $0.04 \times 53.5 = \mathbf{2.14 \text{ g}}$
12. $\text{CaCO}_3(\text{s}) + 2\text{HCl}(\text{g}) \longrightarrow \text{CaCl}_2(\text{aq}) + \text{H}_2\text{O}(\text{g}) + \text{CO}_2(\text{g})$
 Number of mole of $\text{CO}_2 = 1.43 / 44 = 0.0325$
 Number of mole of $\text{CaCO}_3 = 0.0325$
 Mass of pure $\text{CaCO}_3 = 0.0325 \times 100 = 3.25 \text{ g}$
 $\% \text{ by mass of } \text{CaCO}_3 \text{ in the sample} = 3.25 / 3.8 \times 100\% = \mathbf{85.5\%}$
13. Number of mole of $\text{N}_2 = 84 / 28 = 3$
 Number of mole of $\text{NaN}_3 = 3 / 3 \times 2 = 2$
 Mass of $\text{NaN}_3 = 2 \times (23 + 14 \times 3) = \mathbf{130 \text{ g}}$
14. (a) Number of mole of $\text{LiOH} = 50 / (7 + 16 + 1) = 2.08$
 Number of mole of CO_2 absorbed = $2.08 / 2 = 1.04$
 Mass of CO_2 absorbed = $1.04 \times 44 = \mathbf{45.83 \text{ g}}$
- (b) Number of mole of $\text{CO}_2 = 100 / 44 = 2.27$
 Number of mole of $\text{Li}_2\text{CO}_3 = 2.27$
 Mass of $\text{Li}_2\text{CO}_3 = 2.27 \times (7 \times 2 + 12 + 16 \times 3) = \mathbf{167.98 \text{ g}}$
15. (a) $2\text{Ag}_2\text{O}(\text{s}) \longrightarrow 4\text{Ag}(\text{s}) + \text{O}_2(\text{g})$
- (b) Number of mole of $\text{Ag} = 6.52 / 108 = 0.06$
 Number of mole of $\text{Ag}_2\text{O} = 0.06 / 4 \times 2 = 0.03$
 Mass of $\text{Ag}_2\text{O} = 0.03 \times (108 \times 2 + 16) = 6.96 \text{ g}$
 $\% \text{ by mass of } \text{Ag}_2\text{O} = 6.96 / 8 \times 100\% = \mathbf{87\%}$



Number of moles of Hg present = $21.5 / 200.6 = 0.107 \text{ mol}$

Number of moles of Br_2 present = $15.6 / 159.8 = 0.0976 \text{ mol}$

According to the equation, 1 mole of Hg reacts with 1 mole of Br_2 to produce 1 mole of HgBr_2 . During the reaction, 0.0976 mole of Br_2 reacted with 0.0976 mole of Hg. Therefore Hg was in excess. The amount of Br_2 limited the amount of HgBr_2 produced.

Number of moles of HgBr_2 produced = 0.0976 mol

Molar mass of $\text{HgBr}_2 = 200.6 + 2 \times 79.9 = 360.4 \text{ g mol}^{-1}$

(a) Mass of HgBr_2 produced = $0.0976 \times 360.4 = \mathbf{35.2 \text{ g}}$

(b) Mass of Hg reacted = $0.0976 \times 200.6 = 19.6 \text{ g}$

Mass of Hg left = $(21.5 - 19.6) \text{ g} = \mathbf{1.9 \text{ g}}$

17. (a) Number of mole of TiCl_4 present = $4.75 \times 10^7 / 190 = 2.5 \times 10^5$

Number of mole of Mg present = $1.46 \times 10^7 / 24.3 = 6.01 \times 10^5$

\therefore Mg was in excess.

Theoretically, number of mole of Ti produced = 2.5×10^5

Theoretical yield of Ti = $2.5 \times 10^5 \times 48 = \mathbf{1.2 \times 10^7 \text{ g}}$

(b) % yield = $1.06 \times 10^7 / 1.2 \times 10^7 \times 100\% = \mathbf{88.3\%}$

18. (a) Number of mole of Li present = $8.28 / 7 = 1.18$

Number of mole of N_2 present = $10.6 / 28 = 0.37$

\therefore N_2 was in excess.

Theoretically, number of mole of Li_3N produced = $1.18 / 6 \times 2 = 0.39$

Theoretical yield of $\text{Li}_3\text{N} = 0.39 \times (7 \times 3 + 14) = \mathbf{13.65 \text{ g}}$

(b) % yield = $3.97 / 13.65 \times 100\% = \mathbf{29.08\%}$

Suggested Answers on Note (Chapter 10) P.44 – 47

1. Molar mass of $\text{MnO}_2 = 54.9 + 2 \times 16.0 = 86.9 \text{ g mol}^{-1}$
 Mass of MnO_2 in the nodule = $0.0400 \times 86.9 = 3.48 \text{ g}$
 Percentage by mass of MnO_2 in the nodule = $(3.48 / 15.0) \times 100\% = 23.2\%$
2. (a) To prevent the condensed water from running back to the tube and crack the hot glass.
- (b) Test the liquid with dry cobalt(II) chloride paper.
 The liquid turns the paper from blue to pink.
- (c) To prevent 'sucking back' of the liquid.
- (d) Formula mass of $\text{FeSO}_4 \cdot x\text{H}_2\text{O}$
 $= (55.8 + 32.1 + 4 \times 16.0) + x(2 \times 1.0 + 16.0)$
 $= 151.9 + 18x$
- 1 mole of $\text{FeSO}_4 \cdot x\text{H}_2\text{O}$ contains x moles of H_2O .
 i.e. $(151.9 + 18x) \text{ g}$ of $\text{FeSO}_4 \cdot x\text{H}_2\text{O}$ contain $18x \text{ g}$ of H_2O .
- 30.6 g of $\text{FeSO}_4 \cdot x\text{H}_2\text{O}$ contain 13.9 g of H_2O .
- $$18x / (151.9 + 18x) = 13.9 / 30.6 \Rightarrow x = 7$$

3. (a) Suppose we have 100 g of glucose, so there are 40.0 g of carbon, 6.60 g of hydrogen and 53.4 g of oxygen.

	Carbon	Hydrogen	Oxygen
Mass of element in the compound	40.0 g	6.60 g	53.4 g
Relative atomic mass	12.0	1.0	16.0
Number of moles of atoms that combine	$40.0 / 12.0 = 3.33 \text{ mol}$	$6.60 / 1.0 = 6.60 \text{ mol}$	$53.4 / 16.0 = 3.33 \text{ mol}$
Mole ratio of atoms	$3.33 / 3.33 = 1$	$6.60 / 3.33 = 2$	$3.33 / 3.33 = 1$

\therefore the empirical formula of glucose is CH_2O .

- (b) Let $(\text{CH}_2\text{O})_n$ be the molecular formula of glucose.
 Relative molecular mass of glucose = $n(12.0 + 2 \times 1.0 + 16.0) = 30n$
 $\therefore 30n = 180 \Rightarrow n = 6$
 \therefore the molecular formula of glucose is $(\text{CH}_2\text{O})_6$ or $\text{C}_6\text{H}_{12}\text{O}_6$.

4. (a) Method 1



Molar mass of $\text{Fe}(\text{OH})_3 = 55.8 + 3 \times (16.0 + 1.0) = 106.8 \text{ g mol}^{-1}$

Number of moles of $\text{Fe}(\text{OH})_3 = 5.35 / 106.8 = 0.0501 \text{ mol}$

According to the equation, 2 moles of $\text{Fe}(\text{OH})_3$ give 3 moles of H_2O upon heating.

\therefore number of H_2O formed = $3/2 \times 0.0501 = 0.0752 \text{ mol}$

Molar mass of $\text{H}_2\text{O} = 2 \times 1.0 + 16.0 = 18.0 \text{ g mol}^{-1}$

Mass of H_2O formed = $0.0752 \times 18.0 = 1.35 \text{ g}$

Method 2



Molar mass of $\text{Fe}(\text{OH})_3 = 55.8 + 3 \times (16.0 + 1.0) = 106.8 \text{ g mol}^{-1}$

Molar mass of $\text{H}_2\text{O} = 2 \times 1.0 + 16.0 = 18.0 \text{ g mol}^{-1}$

According to the equation, 2 moles of $\text{Fe}(\text{OH})_3$ give 3 moles of H_2O upon heating.

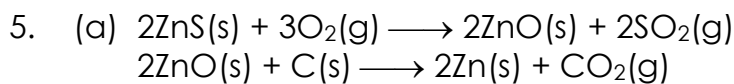
\therefore $2 \times 106.8 \text{ g}$ of $\text{Fe}(\text{OH})_3$ give $3 \times 18.0 \text{ g}$ of H_2O upon heating.

Mass of H_2O formed = $5.35 \times (3 \times 18.0) / (2 \times 106.8) = 1.35 \text{ g}$

(b) Suppose we have 100 g of the oxide, so there are 72.4 g of iron and 27.6 g of oxygen.

	Iron	Oxygen
Mass of element in the oxide	72.4 g	27.6 g
Relative atomic mass	55.8	16.0
Number of moles of atoms that combine	$72.4 / 55.8 = 1.30 \text{ mol}$	$27.6 / 16.0 = 1.73 \text{ mol}$
Mole ratio of atoms	$1.30 / 1.30 = 1.00$	$1.73 / 1.30 = 1.33$
Simplest whole number ratio of atoms	$1 \times 3 = 3$	$1.33 \times 3 = 4$

\therefore the empirical formula of the oxide is Fe_3O_4 .



- (b) Molar mass of ZnO = 65.4 + 16.0 = 81.4 g mol⁻¹
 Number of moles of ZnO = 48.8 / 81.4 = 0.600 mol

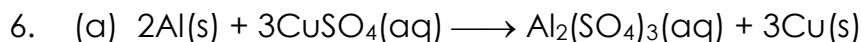
According to the equation, 2 moles of ZnO require 1 mole of C for reduction to give 2 moles of Zn.

$$\therefore \text{ number of moles of Zn obtained} = 0.600 \text{ mol}$$

$$\text{ number of moles of C required} = 0.600 / 2 = 0.300 \text{ mol}$$

(i) Mass of Zn obtained = 0.600 x 65.4 = 39.2 g

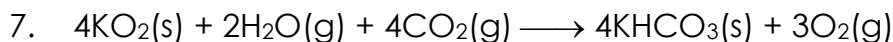
(ii) Mass of C required = 0.300 x 12.0 = 3.60 g



- (b) Number of moles of Al reacted = 1.61 / 27.0 = 0.0596 mol
 According to the equation, 2 moles of Al react to give 3 moles of Cu.
 \therefore number of moles of Cu produced = 3 / 2 x 0.0596 = 0.0894 mol

$$\text{Theoretical yield of Cu} = 0.0894 \times 63.5 = 5.68 \text{ g}$$

$$\text{Percentage yield of Cu} = (2.58 / 5.68) \times 100\% = 45.4\%$$



- Molar mass of CO₂ = 12.0 + 2 x 16.0 = 44.0 g mol⁻¹
 Number of moles of CO₂ exhaled = 14.0 / 44.0 = 0.318 mol

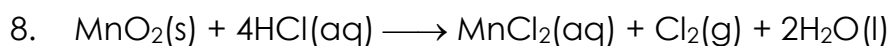
According to the equation, 4 moles of CO₂ require 4 moles of KO₂ for complete reaction.

$$\therefore \text{ number of moles of KO}_2 \text{ required} = 0.318 \text{ mol}$$

(a) Molar mass of KO₂ = 39.1 + 2 x 16.0 = 71.1 g mol⁻¹

$$\text{Theoretical mass of KO}_2 \text{ required} = 0.318 \times 71.1 = 22.6 \text{ g}$$

- (b) Since the process is only 80% efficient, the mass of KO₂ required
 = 22.6 / 80% = 28.3 g



Molar mass of $\text{MnO}_2 = 54.9 + 2 \times 16.0 = 86.9 \text{ g mol}^{-1}$

Number of moles of MnO_2 present = $217 / 86.9 = 2.50 \text{ mol}$

Molar mass of $\text{HCl} = 1.0 + 35.5 = 36.5 \text{ g mol}^{-1}$

Number of moles of HCl present = $274 / 36.5 = 7.51 \text{ mol}$

(a) According to the equation, 1 mole of MnO_2 reacts with 4 moles of HCl to produce 1 mole of Cl_2 . In this case, the mole ratio of MnO_2 to HCl was 1:3. Therefore all the HCl would be used up. The limiting reagent was HCl .

(b) Number of moles of Cl_2 produced = $7.51 / 4 = 1.88 \text{ mol}$

Molar mass of $\text{Cl}_2 = 2 \times 35.5 = 71.0 \text{ g mol}^{-1}$

Mass of Cl_2 produced = $1.88 \times 71.0 = 133 \text{ g}$

(c) Number of moles of MnO_2 used = $7.51 / 4 = 1.88 \text{ mol}$

Mass of MnO_2 used = $1.88 \times 86.9 = 163 \text{ g}$

Mass of MnO_2 (excess reagent) left = $(217 - 163) \text{ g} = 54 \text{ g}$