Quiz (Empirical Formula, Molecular Formula and Structural Formula)

1. Compound *Y* was found to contain iron and oxygen only. Experiments showed that it contains 70% iron and 30% oxygen by mass. Calculate the empirical formula of *Y*.
(Relative atomic masses: O = 16.0, Fe = 55.8)

2. An experiment was performed to determine the empirical formula of an oxide of magnesium. The experimental results are tabulated below.

Item	Mass (g)
Crucible + lid	28.092
Crucible + lid + magnesium	28.698
Crucible + lid + oxide of magnesium	29.103

Determine the empirical formula of the oxide of magnesium using the above data.
(Relative atomic masses: O = 16.0, Mg = 24.3)

3. 1.200 g of a compound containing only carbon, hydrogen and oxygen gave 1.173 g of carbon dioxide and 0.240 g of water on complete combustion. Find the empirical formula of the compound.
(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

4. A compound has the empirical formula C_xH_y. On analysis, 1.000 g of the compound was found to contain 0.857 g of carbon. Find the values of *x* and *y*.
(Relative atomic masses: H = 1.0, C = 12.0)

5. Compound *X* contains 26.95% sulphur, 13.44% oxygen and 59.61% chlorine by mass. Find the empirical formula of *X*.
(Relative atomic masses: O = 16.0, S = 32.1, Cl = 35.5)

6. A compound has an empirical formula CH_2 and a relative molecular mass of 42.0. Determine its molecular formula.
(Relative atomic masses: H = 1.0, C = 12.0)

7. Compound *X* was found to contain carbon and hydrogen only. Experiments showed that it contained 80% carbon and 20% hydrogen by mass. If its relative molecular mass is 30.0, calculate the empirical formula and molecular formula of *X*.
(Relative atomic masses: H = 1.0, C = 12.0)

8. Compound *Z* containing only carbon, hydrogen and oxygen burnt completely in air to form carbon dioxide and water as the only products. 2.43 g of *Z* gave 3.96 g of carbon dioxide and 1.35 g of water.
Determine the empirical formula of *Z*. If its relative molecular mass is 162.0, determine the molecular formula of *Z*.
(Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

9. 5.60 g of hydrated copper(II) sulphate $CuSO_4 \bullet nH_2O$ was heated in a crucible to drive off the water of crystallization. The white residue was anhydrous copper(II) sulphate, which was found to have a mass of 3.59 g.
   (a) Deduce a reasonable value for n.
   (b) Explain why the answer you gave in (a) differs a bit from the value actually calculated.
   (Relative atomic masses: H = 1.0, O = 16.0, S = 32.1, Cu = 63.5)

10. A compound containing only carbon, hydrogen and oxygen. 0.81 g of the compound gave 1.32 g of carbon dioxide and 0.45 g of water on complete combustion. Find the empirical formula of the compound. If the relative molecular mass of the compound is 320.0, find its molecular formula.
    (Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

11. A compound was found to contain 40.00% by mass of carbon, 6.67% by mass of hydrogen and 53.33% by mass of oxygen. It has a relative molecular mass of 60.0. Calculate its molecular formula.
    (Relative atomic masses: H = 1.0, C = 12.0, O = 16.0)

12. Epsom salts are used as bath salts to relieve aches and pains. They are hydrated salts of magnesium sulphate with formula $MgSO_4 \bullet nH_2O$.
    Experiments were carried out to find the formula of the salt. It was found that it contained 51.22% by mass of water of crystallization. Find the value of n.
    (Relative atomic masses: H = 1.0, O = 16.0, Mg = 24.3, S = 32.1)

Suggested Answer

1. Assume that there are 100 g of *Y*. Then there are 70 g of iron and 30 g of oxygen.

	Fe	**O**
Mass (g)	70	30
Number of moles of atoms (mol)	70 / 55.8 = 1.25	30 / 16.0 = 1.88
Simplest whole number mole ratio of atoms (divided by the smallest number of moles)	1.25 / 1.25 = 1	1.88 / 1.25 = 1.5
multiplied by the smallest possible whole number (2 here) to turn all the values into whole numbers	1 × 2 = 2	1.5 × 2 = 3

∴ the empirical formula of *Y* is Fe_2O_3.

2.

	Mg	**O**
Mass (g)	28.698 – 28.092 = 0.606	29.103 – 28.698 = 0.405
Relative atomic mass	24.3	16.0
Number of moles of atoms (mol)	0.606 / 24.3 = 0.0249	0.405 / 16.0 = 0.0253
Simplest whole number mole ratio of atoms	0.0249 / 0.0249 = 1	0.0253 / 0.0249 = 1.02 ≈ 1

∴ the empirical formula of the oxide of magnesium is MgO.

3. Mass of C in the compound = 1.173 × (12.0 / 12.0 + 16.0 × 2) g = 0.320 g
Mass of H in the compound = 0.240 × (1.0 × 2 / 1.0 × 2 + 16.0) g = 0.0267 g
Mass of O in the compound = (1.200 – 0.320 – 0.0267) g = 0.853 g

	C	**H**	**O**
Mass (g)	0.320	0.0267	0.853
Relative atomic mass	12.0	1.0	16.0
Number of moles of atoms (mol)	0.320 / 12.0 = 0.0267	0.0267 / 1.0 = 0.0267	0.853 / 16.0 = 0.0533
Simplest whole number mole ratio of atoms	0.0267 / 0.0267 = 1	0.0267 / 0.0267 = 1	0.0533 / 0.0267 = 2

∴ the empirical formula of the compound is CHO_2.

4.

	C	**H**
Mass (g)	0.857	1.000 – 0.857 = 0.143
Relative atomic mass	12.0	1.0
Number of moles of atoms (mol)	0.857 / 12.0 = 0.0714	0.143 / 1.0 = 0.143
Simplest whole number mole ratio of atoms	0.0714 / 0.0714 = 1	0.143 / 0.0714 = 2

∴ the empirical formula of the compound is CH_2.

5. Assume that there are 100 g of *X*. Then, there are 26.95 g of sulphur, 13.44 g of oxygen and 59.61 g of chlorine.

	S	O	Cl
Mass (g)	26.95	13.44	59.61
Relative atomic mass	32.1	16.0	35.5
Number of moles of atoms (mol)	26.95 / 32.1 = 0.840	13.44 / 16.0 = 0.84	59.61 / 35.5 = 1.68
Simplest whole number mole ratio of atoms	0.840 / 0.840 = 1	0.84 / 0.84 = 1	1.68 / 0.84 = 2

∴ the empirical formula of the compound is $SOCl_2$.

6. Let the molecular formula of the compound be $(CH_2)_n$, where n is a whole number.

Relative molecular mass of $(CH_2)_n$ = 42.0
$n\,(12.0 + 1.0 \times 2) = 42.0$
$\Rightarrow$ $n = 3$

∴ the molecular formula of the compound is $(CH_2)_3$, i.e. C_3H_6.

7. Assume that there are 100 g of *X*. Then there are 80 g of carbon and 20 g of hydrogen.

	C	**H**
Mass (g)	80	20
Relative atomic mass	12.0	1.0
Number of moles of atoms (mol)	80 / 12.0 = 6.67	20 / 1.0 = 20
Simplest whole number mole ratio of atoms	6.67 / 6.67 = 1	20 / 6.67 = 3

∴ the empirical formula of *X* is CH_3.

Let the molecular formula of *X* be $(CH_3)_n$, where n is the whole number.

Relative molecular mass of $(CH_3)_n$ = 30.0
n (12.0 + 1.0 × 3) = 30.0
15.0 n = 30.0
⇒ n = 2

∴ the molecular formula of *X* is C_2H_6.

Note: 2.99 can be rounded off to 3, but 2.8 is usually NOT rounded off to 3.

8. Since all the C in CO_2 and H in H_2O came from *Z*,

mass of C in *Z* = 3.96 g × (12.0 / 12.0 + 16.0 × 2) = 1.08 g;
mass of H in *Z* = 1.35 g × (1.0 × 2 / 1.0 × 2 + 16.0) = 0.15 g

The rest of mass of *Z* must come from oxygen.

∴ mass of O in *Z* = (2.43 – 1.08 – 0.15) g = 1.20 g

Now go on to find the empirical formula of *Z* as follows:

	C	**H**	**O**
Mass (g)	1.08	0.15	1.20
Relative atomic mass	12.0	1.0	16.0
Number of moles of atoms (mol)	1.08 / 12.0 = 0.090	0.15 / 1.0 = 0.15	1.20 / 16.0 = 0.075
Simplest whole number mole ratio of atoms	0.090 / 0.075 = 1.2	0.15 / 0.075 = 2	0.075 / 0.075 = 1
multiplied by the smallest possible whole number (5 here) to turn all values into whole numbers	1.2 × 5 = 6	2 × 5 = 10	1 × 5 = 5

∴ the empirical formula of *Z* is $C_6H_{10}O_5$.

Let the molecular formula of *Z* be $(C_6H_{10}O_5)_n$, where n is the whole number.

Relative molecular mass of $(C_6H_{10}O_5)_n$ = 162.0
n (12.0 × 6 + 1.0 × 10 + 16.0 × 5) = 162.0
162.0 n = 162.0
⇒ n = 1

∴ the molecular formula of *Z* is $C_6H_{10}O_5$.

9. (a) Mass of water of crystallization = (5.60 – 3.59) g = 2.01 g

	$CuSO_4$	H_2O
Mass (g)	3.59	2.01
Formula mass	159.6	18.0
Number of moles of formula units (mol)	3.59 / 159.6 = 0.0225	2.01 / 18.0 = 0.112
Simplest whole number mole ratio of formula units	0.0225 / 0.0225 = 1	0.112 / 0.0225 = 4.98

Since n should be a whole number, a reasonable value of n would be 5.

(b) The experimental value of n (4.98) is lower than 5. This might be due to two reasons:
(1) Not all water of crystallization has been removed in the heating process.
(2) The anhydrous salt has absorbed some moisture from the atmosphere during weighing.

10. Mass of C in the compound = 1.32 × 12.0
12.0 + 16.0 × 2
g = 0.36 g
Mass of H in the compound = 0.45 × 1.0 × 2
1.0 × 2 + 16.0
g = 0.05 g
Mass of O in the compound = (0.81 – 0.36 – 0.05) g = 0.40 g
the empirical formula of the compound is C6H10O5.
Let the molecular formula of the compound be (C6H10O5)n.
320.0 = n × (12.0 × 6 + 1.0 × 10 + 16.0 × 5)
n = 1.98 2
the molecular formula of the compound is C12H20O10.

11. Assume that there are 100 g of the compound. Then, there are 40.00 g
of carbon, 6.67 g of hydrogen and 53.33 g of oxygen.
the empirical formula of the compound is CH2O.
Let the molecular formula of the compound be (CH2O)n.
60.0 = n × (12.0 + 1.0 × 2 + 16.0)
n = 2
the molecular formula of the compound is C2H4O2.

12. Assume that there are 100 g of Epsom salt. Then, there are 51.22 g
of water of crystallization and (100 – 51.22) g = 48.78 g of MgSO4.
the value of n is 7.
MgSO4 H2O
Mass (g) 48.78 51.22
Formula mass 24.3 + 32.1 + 16.0 × 4 = 120.4 1.0 × 2 + 16.0 = 18.0
Number of moles of
formula units (mol)
48.78
120.4

= 0.4051 51.22
18.0
= 2.85
Simplest whole
number mole ratio of
formula units
0.4051
0.4051
= 1 2.85
0.4051
= 7.04 7
C H O
Mass (g) 0.36 0.05 0.40
Relative
atomic mass
12.0 1.0 16.0
Number of
moles of
atoms (mol)
0.36
12.0
= 0.03 0.05
1.0
= 0.05 0.40
16.0
= 0.025
Simplest
whole
number mole
ratio of atoms
0.03
0.025
= 1.2 0.05
0.025
= 2 0.025
0.025
= 1
1.2 × 5 = 6 2 × 5 = 10 1 × 5 = 5
C H O
Mass (g) 40.00 6.67 53.33
Relative atomic
mass
12.0 1.0 16.0
Number of moles
of atoms (mol)
40.00
12.0
= 3.33 6.67
1.0

= 6.67 53.33
16.0
= 3.33
Simplest whole
number mole
ratio of atoms
3.33
3.33
= 1 6.67
3.33
= 2 3.33
3.33
= 1