Year 11 Chemistry – Empirical Formula NOTES

Empirical Formula is the simplest whole number ratio of elements in a compound. For example, the empirical formula of ethane, C_2H_6, would be CH_3.

You will encounter three types of questions with respect to empirical formula.

1. Percentage Composition
2. Hydrated Compounds
3. Combustion Analysis

<u>Percentage composition</u>

The percentage composition of an element on a compound is the % by mass of the element.

Example 1. Determine the percentage composition of carbon in carbon dioxide.

% C = M (C) / M (CO_2) x 100 = 12.01 / 44.01 x 100 = 27.28%

Example 2. Determine the empirical formula of a compound containing 60% S and 40% oxygen.

1. If % composition is given; assume a mass of 100 g of compound.
2. Find the moles present of each element.
3. Find the ratio of moles of each element, to deduce the empirical formula

Therefore,

1. m (S) = 60 g and m (O) = 40 g

2. n (S) = 60 g / 32.06 g mol^{-1} = 1.87 mol
   n (O) = 40 g / 16 g mol^{-1} = 2.5 mol

3. Ratio

O	S
2.5 mol / 1.87 mol	1.87 mol / 1.87 mol
= 1.33	= 1

Ratio is: $SO_{1.33}$

Multiply each element by 3 to form a whole number ratio

Empirical Formula = S_3O_4

Hydrated Compounds

Some compounds crystallise with water molecules in their structure. For example, $BaCl_2 \cdot xH_2O$.

Your task is to determine the value of x, and also the empirical formula, using the provided information.

A 5.00 g sample of hydrated barium chloride is heated to drive of the water molecules. After heating, 4.26 g of anhydrous barium chloride ($BaCl_2$) remains. Determine the value of x, and the empirical formula.

1. Find the moles of water present.

$n(H_2O) = m / M = (5.00\ g – 4.26\ g) / 18.016\ g\ mol^{-1} = 0.041\ mol$

2. Find the moles of anhydrous compound.

$n(BaCl_2) = m / M = 4.26\ g / 208.2 = 0.020\ mol$

3. Find mole ratio of components.

Ratio	H_2O	$BaCl_2$
	0.041 mol / 0.020 mol	0.020 mol / 0.020 mol
	= 2	= 1

x = 2

Empirical Formula = $BaCl_2 \cdot 2H_2O$

Combustion Analysis

Hydrocarbons can be combusted in oxygen following the general equation:

$C_xH_y + O_2 \rightarrow CO_2 + H_2O$

Example 1. A 5.000 g sample of an unknown hydrocarbon was combusted in excess oxygen to produce 14.97 g of carbon dioxide and 8.143 g of water vapour. Determine the empirical formula of the unknown hydrocarbon.

C_xH_y	+	O_2	→	CO_2	+	H_2O
5.00 g				14.97 g		8.143 g

Firstly, all the moles of carbon in this equation are present in the hydrocarbon and then the carbon dioxide. All of the moles of hydrogen in this equation are present in the hydrocarbon and then the water. Thus, if the moles of carbon and hydrogen are determined in the products, the ratio is able to be determined in the reactants.

1. Find moles of CO_2, and C.

$n(CO_2) = 14.97 \text{ g} / 44.01 \text{ g mol}^{-1} = 0.34 \text{ mol}$
$n(C) = n(CO_2) = 0.34 \text{ mol}$

2. Find moles of H_2O, and H.

$n(H_2O) = 8.143 \text{ g} / 18.016 \text{ g mol}^{-1} = 0.452 \text{ mol}$
$n(H) = 2 \times n(H_2O) = 0.904 \text{ mol}$

3. Find ratio of moles and empirical formula.

Ratio	C	H
	0.34 mol / 0.34 mol	0.904 mol / 0.34 mol
	= 1	= 2.66

Ratio is: $CH_{2.66}$

Multiply each element by 3 to form a whole number ratio

Empirical Formula = C_3H_8

Example 2. The molecular weight of this unknown hydrocarbon was determined to be 88.19 g mol^{-1}. Identify the molecular formula.

Absolute Classic #20

A 1.98 g sample of cobalt (II) chloride hydrate is heated over a burner. When cooled, the mass of the remaining dehydrated compound is found to be 1.55 g. Determine the empirical formula of the original hydrate. How can the scientist ensure that all of the water is removed from the sample?

ANSWER

1. Find the moles of water present.

$n(H_2O) = m / M = (1.98\ g – 1.55\ g) / 18.016\ g\ mol^{-1} = 0.0239\ mol$

2. Find the moles of anhydrous compound.

$n(CoCl_2) = m / M = 1.55\ g / 129.84\ g\ mol^{-1} = 0.01194\ mol$

3. Find mole ratio of components.

Ratio	H_2O	$CoCl_2$
	0.0239 mol / 0.01194 mol = 2	0.01194 mol / 0.01194 mol = 1

x = 2

Empirical Formula = $CoCl_2 \cdot 2H_2O$