Unit 6

On completion of the unit you should be able to:

1. balance chemical equations.
2. interpret balanced equations in terms of moles, mass units and number of particles.
3. given the number of moles (or mass or number of particles) of one substance calculate the number of moles (or mass or number of particles) of another substance in the balanced equation.
4. define limiting and excess reactants.
5. identify the following types of reactions:
   - synthesis
   - decomposition
   - combustion
   - single-replacement
   - double-replacement
   - neutralization
6. define endothermic and exothermic reactions.
7. define acids and bases.

Recognize these types of reactions

Write this type of reaction

Today's focus.

6.5 Types of Reactions

Five major classes of chemical reactions

1. Combination or Synthesis reactions
2. Decomposition Reactions
3. Combustion reactions

6.6 Endothermic and exothermic reactions

Bill Nye The Science Guy on Chemical Reactions

Learn to *recognize* Synthesis reactions

- synthesis
- decomposition
- combustion
- single-replacement
- double-replacement
- neutralization

Combination or Synthesis reactions

Combination or synthesis reactions include reactions of the general form:

$A + B \longrightarrow AB$

where a single more complex compound is formed from the reaction of two or more substances.

Example

Synthesis of ammonia: $N_2\ (g) + 3\ H_2\ (g) \longrightarrow 2\ NH_3\ (g)$

Synthesis of water: $2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$

Synthesis of potassium chloride: $2\ K\ (s) + Cl_2\ (g) \rightarrow 2\ KCl\ (s)$

Synthesis of iron (III) oxide (rust): $4\ Fe\ (s) + 3\ O_2\ (g) \rightarrow 2\ Fe_2O_3\ (s)$

Learn to *recognize* Decomposition reactions

- ~~synthesis~~
- decomposition
- combustion
- single-replacement
- double-replacement
- neutralization

Decomposition reactions

(Redirected from Decomposition Reactions)

Decomposition reactions include reactions of the general form:

$AB \longrightarrow A + B$

where a single compound breaks down into two or more simpler substances.

Examples

1. The decomposition of potassium chlorate, $KClO_3$.

   $2\ KClO_3\ (s) \longrightarrow 2\ KCl\ (s) + 3\ O_2\ (g)$

2. The decomposition of water, H_2O.

   $2\ H_2O\ (l) \longrightarrow 2\ H_2\ (g) + O_2\ (g)$

Decomposition of magnesium chloride: $MgCl_2\ (s) \rightarrow Mg\ (s) + Cl_2\ (g)$

Decomposition of iron (II) sulfide: $FeS\ (s) \rightarrow Fe\ (s) + S\ (s)$

Learn to *write* Combustion reactions

- synthesis
- decomposition
- **combustion**
- single-replacement
- double-replacement
- neutralization

Combustion reaction involves O_2 (g) as a reactant.

In this course, we will limit combustion to mean the reaction of oxygen with:

- compounds containing carbon and hydrogen.

examples: methane, CH_4; ethane, C_2H_6; propane, C_3H_8; ← gases

- compounds containing carbon, hydrogen, and oxygen.

examples: methyl alcohol, CH_3OH; ethyl alcohol, C_2H_5OH; propyl alcohol, C_3H_7OH; ← liquids

examples: sucrose, $C_{12}H_{22}O_{11}$; glucose, $C_6H_{12}O_6$; ← solids

Do NOT need to memorize these chemical formulae.

Need to know that the products of combustion of this type of compounds are *always*:

CO_2 (g) and H_2O (l)

Examples (unbalanced reactions):

CH_4 (g) + O_2 (g) → CO_2 (g) + H_2O (l)
C_2H_6 (g) + O_2 (g) → CO_2 (g) + H_2O (l)
$C_6H_{12}O_6$ (s) + O_2 (g) → CO_2 (g) + H_2O (l)
C_2H_5OH (l) + O_2 (g) → CO_2 (g) + H_2O (l)

Balanced reactions:

CH_4 (g) + 2 O_2 (g) → CO_2 (g) + 2 H_2O (l)
2 C_2H_6 (g) + 7 O_2 (g) → 4 CO_2 (g) + 6 H_2O (l)
$C_6H_{12}O_6$ (s) + 6 O_2 (g) → 6 CO_2 (g) + 6 H_2O (l)
C_2H_5OH (l) + 3 O_2 (g) → 2 CO_2 (g) + 3 H_2O (l)

BEFORE: From mass percent → Empirical Formula & Chemical Formula

NOW: From combustion analysis of C, H, O containing compounds → Empirical Formula & Chemical Formula

Maple TA type Problem in Assignment 7

Question:

Compounds containing Carbon-Hydrogen-Oxygen can be analysed by burning a sample of the compound in the presence of excess oxygen gas. As a result of the combustion of the compound, all of the carbon in the compound is converted to carbon dioxide gas and all of the hydrogen in the compound is converted to water vapour.

(a) Combustion of 0.4720 g of the compound produced 0.96502 g CO_2 and 0.29627 g H_2O. What is the empirical formula of the compound?

[Hint: Combustion involves O_2 as a reactant. The oxygen in CO_2 and H_2O contains oxygen from the sample and the oxygen from the air. You need to subtract off the oxygen from the air.]

- C_3H_6O
- C_2H_3O
- C_2H_6O

Unbalanced reaction:

$$C_xH_yO_z + O_2\,(g) \rightarrow CO_2\,(g) + H_2O\,(l)$$

(b) The molar mass of the compound is 172.18 g/mole. What is the chemical formula of the compound?

- $C_6H_{12}O_2$
- $C_4H_{12}O_2$
- $C_8H_{12}O_4$

Plan of Action: The plan is to determine the mass of C, H, and O.

From Analysis → Empirical Formula → Chemical Formula Problem

$C_xH_yO_z$

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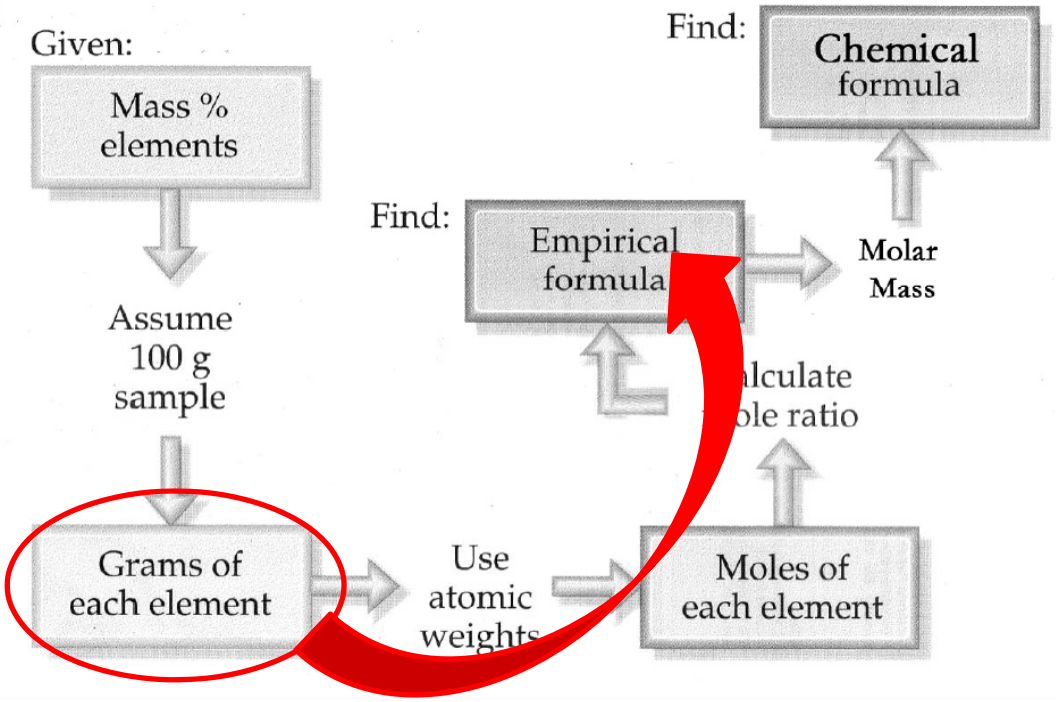
- C_3H_6O
- C_2H_3O
- C_2H_6O

T-26 **Figure 3.11 Procedure for Calculating Empirical Formulas**



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- C_3H_6O
- C_2H_3O
- C_2H_6O

Unbalanced reaction:

$$C_xH_yO_z \rightarrow CO_2\,(g) + H_2O\,(l)$$

0.4720 g (Mass of Compound)

0.96502 g → Moles of CO_2 → Moles of C → **Step 1. Grams of C**

0.29627 g → Moles of H_2O → Moles of H → **Step 2: Grams of H**

To get Grams of O:

Step 3:

Grams of O = Mass of Compound - (Grams of C + Grams of H)

(i.e. This excludes contribution of O from the air.)

From Analysis → Empirical Formula → Chemical Formula Problem

Unbalanced reaction:

$$C_xH_yO_z + O_2\,(g) \rightarrow CO_2\,(g) \quad + \quad H_2O\,(l)$$

0.4720 g	0.96502 g	0.29627 g
	↓	↓
	Moles of CO_2	Moles of H_2O
	↓	↓
	Moles of C	Moles of H
	↓	↓
	Grams of C	Grams of H

3. Grams of Oxygen:

Mass of O = 0.4720 g - **0.2966 g** = **0.1754 g O**

(Mass of Sample)

Mass of C + H = 0.2634g + 0.03315g = **0.2966 g C&H**

1. From mass of CO_2 to Grams of Carbon:

$$0.96502\ \text{g} \cdot \frac{1\ \text{mole}}{44.009\ \text{g}} = 0.02193\ \text{moles}\ CO_2 \cdot \frac{1\ \text{mole C}}{1\ \text{mole}\ CO_2} = 0.02193\ \text{moles C}$$

$$= 0.02193\ \text{moles C} \cdot 12.011\ {}^{g}/_{\text{mole}} = 0.2634\ \text{g C}$$

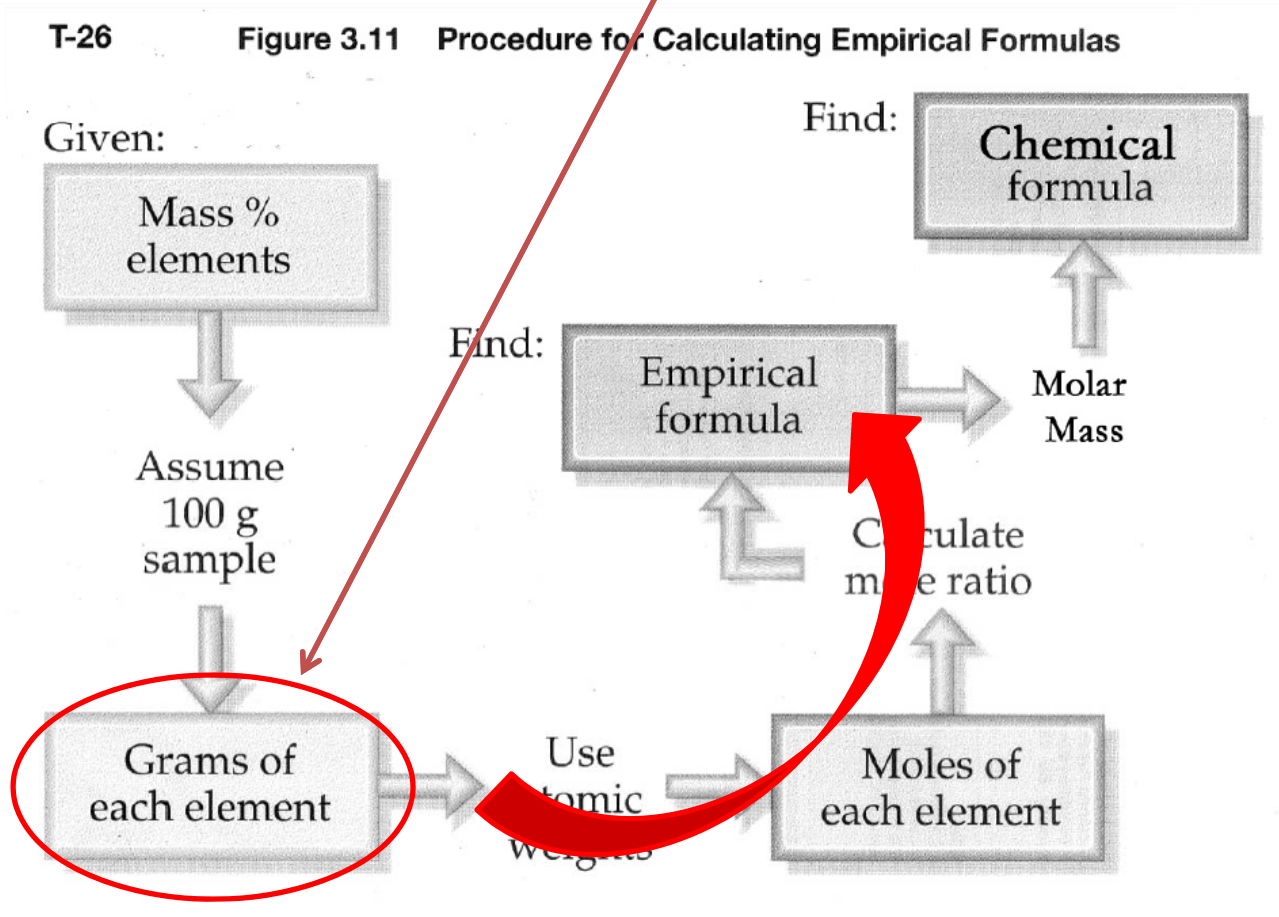
2. From mass of H_2O to Grams of Hydrogen:

$$0.29627\ \text{g} \cdot \frac{1\ \text{mole}}{18.015\ \text{g}} = 0.01645\ \text{moles}\ H_2O \cdot \frac{2\ \text{mole H}}{1\ \text{mole}\ H_2O} = 0.03289\ \text{moles H}$$

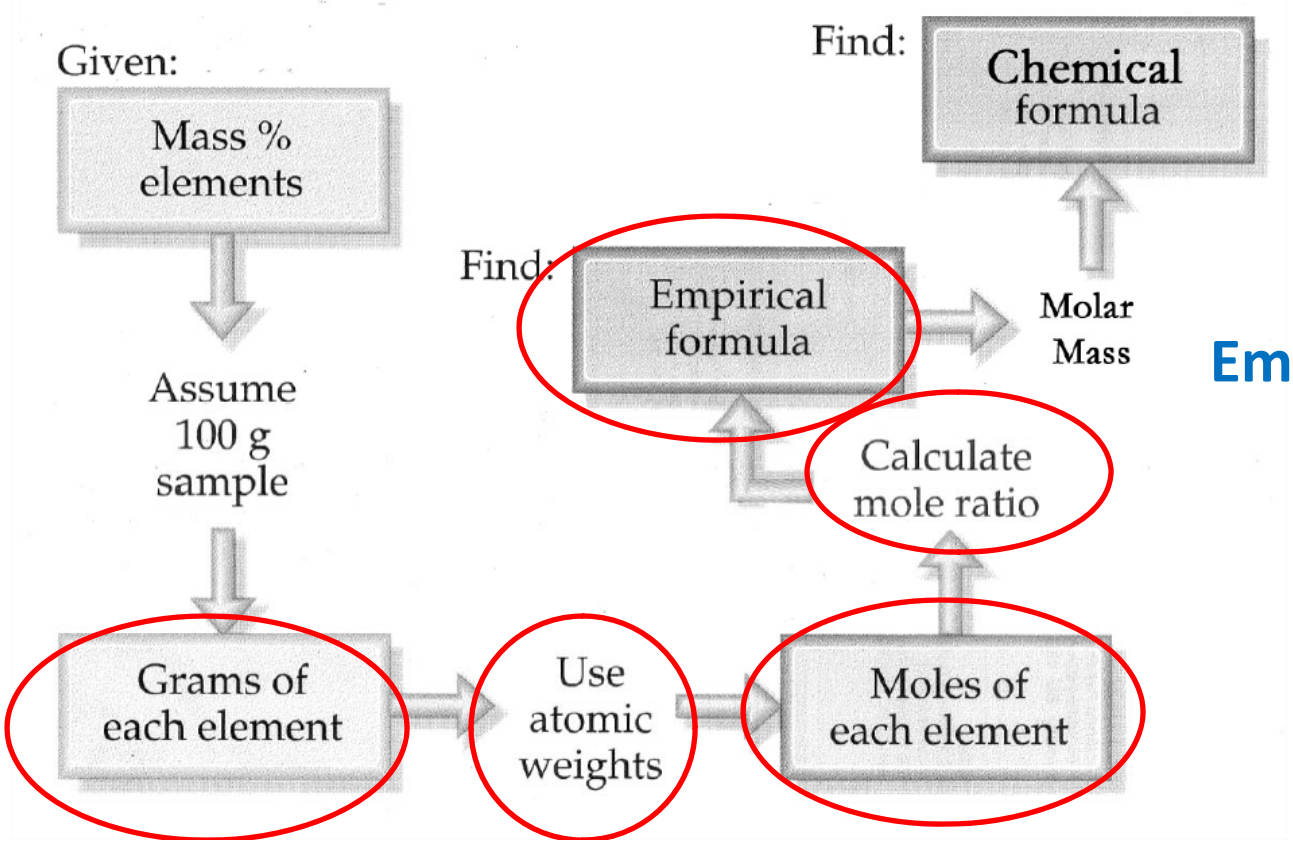
$$= 0.03289\ \text{moles H} \cdot 1.008\ {}^{g}/_{\text{mole}} = 0.03315\ \text{g H}$$

(a) Combustion of 0.4720 g of the compound produced 0.96502 g CO_2 and 0.29627 g H_2O. What is the empirical formula of the compound?

Unknown Compound	Mass of C	Mass of H	Mass of O
$C_xH_yO_z$	0.2634 g C	0.03315 g H	0.1754 g O



	Mass		Moles	(divide by smallest #)	Simplest Mole Ratio
C	0.2634 g C	$\frac{0.2634\ \text{g C}}{12.011\ \frac{\text{g}}{\text{mole}}}$	0.02193 moles C	$\frac{0.02193}{0.01096}$	2
H	0.03315 g H	$\frac{0.03315\ \text{g H}}{1.008\ \frac{\text{g}}{\text{mole}}}$	0.03289 moles H	$\frac{0.03289}{0.01096}$	3
O	**0.1754 g O**	$\frac{0.1754\ \text{g O}}{15.999\ \frac{\text{g}}{\text{mole}}}$	0.01096 moles O	$\frac{0.01096}{0.01096}$	1



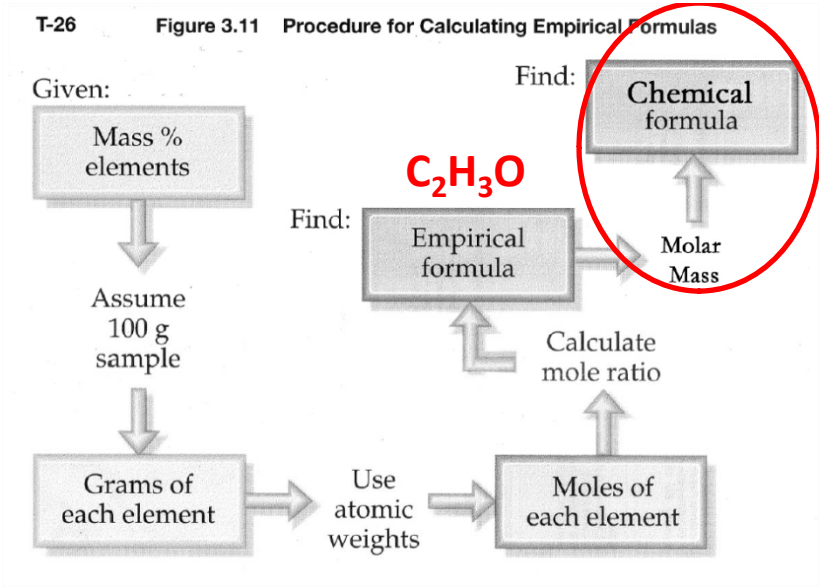
Empirical Formula: C_2H_3O

Empirical formula mass = 43.045 $^{g}/_{mole}$

Find the chemical formula:

(b) The molar mass of the compound is 172.18 $^{g}/_{mole}$.
What is the chemical formula of the compound?

- $C_6H_{12}O_2$
- $C_4H_{12}O_2$
- $C_8H_{12}O_4$



Empirical Formula: C_2H_3O

Empirical formula mass = 43.045 $^{g}/_{mole}$

$$Multiple = \frac{\text{molar mass of compound}}{\text{empirical formula mass}} = \frac{172.18\ \frac{g}{mole}}{43.045\ \frac{g}{mole}} = 4$$

Chemical formula of the compound	C:H:O Ratio	Empirical Formula	*Multiple*
$C_8H_{12}O_4$	2:3:1 x4	C_2H_3O	4

Try some practice problems in Maple TA.

Use Dimensional Analysis in ALL your calculations!!