Mass Percent and Formulas

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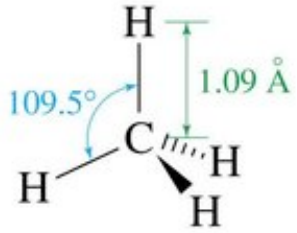
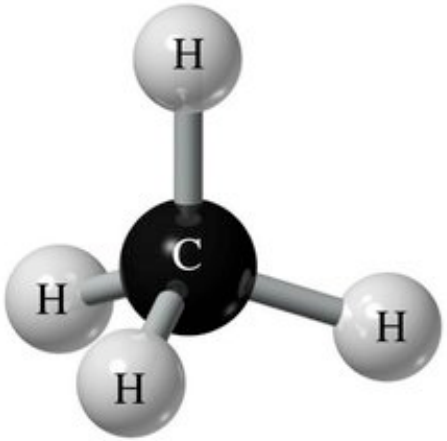
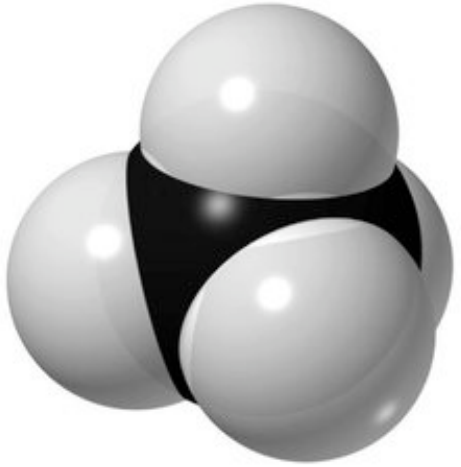
Introduction to mass percent

Imagine a single molecule of methane: CH_4



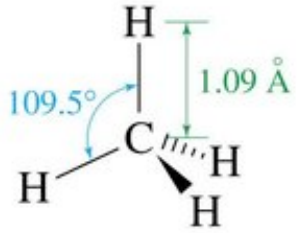


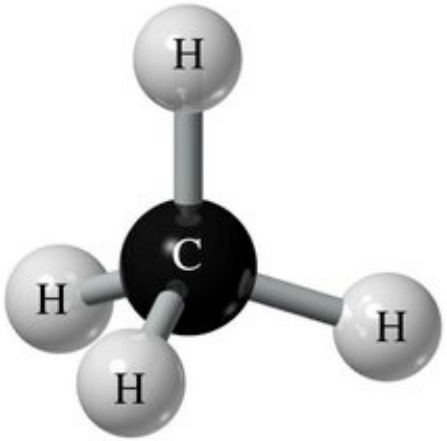


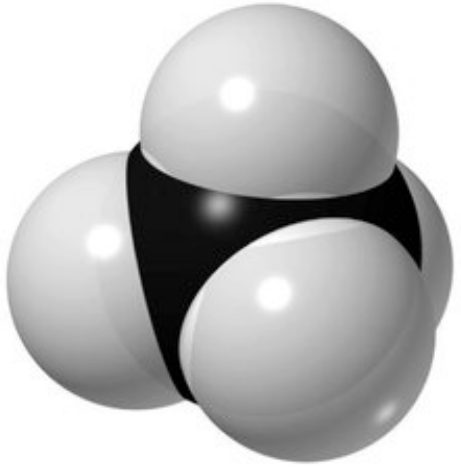
- Ask yourself: *Does this molecule contain more hydrogen or more carbon?*
  - IT DEPENDS!
  - Technically, 4 out of 5 atoms are hydrogen (80%), but….

Introduction to mass percent

CHEMISTS CARE ABOUT MASS PERCENT!







$$\% \text{ Mass} = \frac{\text{mass part}}{\text{mass whole}} \times 100\%$$

How do I calculate the mass percentages for CH_4?

- Remember that the molar mass of CH_4 is 16.04 g/mol:

$$1\ \text{mol}\ CH_4 = 1\ \text{mol C} + 4\ \text{mol H}$$
$$= 1\ (12.01\ \text{g}) + 4\ (1.008\ \text{g}) = 16.04\ \text{g}$$

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$$\%\ \text{C} \rightarrow \frac{1(12.01)\ \text{g}}{16.04\ \text{g}} \times 100\% = 74.90\%\ \text{C}$$

$$\%\ \text{H} \rightarrow \frac{4(1.008)\ \text{g}}{16.04\ \text{g}} \times 100\% = 25.10\%\ \text{H}$$

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SAME THING!

TAKE-HOME MESSAGE

Percent composition is independent of the starting amount!

This is why we usually *assume* we have 100 g or 1 mol.

These are just super easy numbers to work with.

Note: If you wanted to use a strange amount, like 0.27 mol or 74.5 g of substance, your answers would be the same but the math isn't as convenient. BUT you'll still be right. ☺

What is the mass percent of oxygen in each substance?

H_2O_2 versus H_2O

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H_2O_2 (hydrogen peroxide)

The molar mass of H_2O_2 is 34.02 g/mol.

H_2O (water)

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$$\frac{2(1.008)\ \text{g}}{34.02\ \text{g}} \times 100\% = 5.93\%\ \text{H}$$

$$\frac{2(1.008)\ \text{g}}{18.02\ \text{g}} \times 100\% = 11.19\%\ \text{H}$$

Note: The numbers will not always be exactly 100%.

Most often, we use mass percentages to help us figure out what compound we have.

These are called EMPIRICAL FORMULAS.

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Q: Is our formula then $N_{30.4}O_{69.6}$?

A: No! Why? A chemical formula represents number of atoms in a compound, not the mass of each.

We must convert the masses to *moles*.

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$$\text{N} \rightarrow 30.4\ \text{g N} \times \frac{1\ \text{mol N}}{14.01\ \text{g N}} = 2.17\ \text{mol N}$$

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Q: So, is our formula then $N_{2.17}O_{4.35}$?

A: No! Why? Atoms can't be fractional.

We need a whole number ratio!

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Divide the number of moles by the SMALLEST value!

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THIS IS IT!

Our compound has the empirical formula: NO_2 (nitrogen dioxide)

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This is the same as NO_2!

Q: How do we differentiate between NO_2 and N_2O_4?

A: You can use the molar masses of NO_2 and N_2O_4

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$$\text{O} \rightarrow 36.92\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 2.31\ \text{mol O}$$

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Uh... these aren't whole numbers!

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Multiply to get whole numbers!

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Al_2O_3

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Assuming a 100 g sample of cisplatin ($Pt_aN_bH_cCl_d$):

$$\text{Pt} \rightarrow 65.02\ \text{g Pt} \times \frac{1\ \text{mol Pt}}{195.1\ \text{g Pt}} = 0.3333\ \text{mol Pt} \rightarrow \frac{0.3333\ \text{mol Pt}}{0.3333} = 1\ \text{Pt}$$

$$\text{N} \rightarrow 9.34\ \text{g N} \times \frac{1\ \text{mol N}}{14.01\ \text{g N}} = 0.667\ \text{mol N} \rightarrow \frac{0.667\ \text{mol N}}{0.3333} = 2\ \text{N}$$

$$\text{H} \rightarrow 2.02\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 2.00\ \text{mol H} \rightarrow \frac{2.00\ \text{mol H}}{0.3333} = 6\ \text{H}$$

$$\text{Cl} \rightarrow 23.63\ \text{g Cl} \times \frac{1\ \text{mol Cl}}{35.45\ \text{g Cl}} = 0.6666\ \text{mol Cl} \rightarrow \frac{0.6666\ \text{mol Cl}}{0.3333} = 2\ \text{Cl}$$

$PtN_2H_6Cl_2$

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Assuming a 100 g sample of the halohydrocarbon ($Cl_aC_bH_c$):

$$\text{Cl} \rightarrow 71.65\ \text{g Cl} \times \frac{1\ \text{mol Cl}}{35.45\ \text{g Cl}} = 2.021\ \text{mol Cl} \rightarrow \frac{2.021\ \text{mol Cl}}{2.021} = 1\ \text{Cl}$$

$$\text{C} \rightarrow 24.27\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 2.021\ \text{mol C} \rightarrow \frac{2.021\ \text{mol C}}{2.021} = 1\ \text{C}$$

$$\text{H} \rightarrow 4.07\ \text{g H} \times \frac{1\ \text{mol H}}{1.008\ \text{g H}} = 4.04\ \text{mol H} \rightarrow \frac{4.04\ \text{mol H}}{2.021} = 2\ \text{H}$$

The empirical formula is $ClCH_2$.

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$$n = \frac{Molar\ mass}{Empirical\ formula\ mass} = \frac{98.96\ \text{g}}{49.48\ \text{g}} = 2$$

The molecular formula is $(ClCH_2)_2$ or $Cl_2C_2H_4$.

For each of the following, the molecular formula is given. Determine the empirical formula for each compound.

H_2O_2

C_6H_6

C_4H_{10}

CCl_4

$C_2H_4Cl_4O_2$

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C_6H_6	78.11 g/mol
C_4H_{10}	58.12 g/mol
CCl_4	153.81 g/mol
$C_2H_4Cl_4O_2$	201.85 g/mol

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C_4H_{10}	58.12 g/mol	C_2H_5	29.06 g/mol
CCl_4	153.81 g/mol	CCl_4	153.81 g/mol
$C_2H_4Cl_4O_2$	201.85 g/mol	CH_2Cl_2O	100.93 g/mol

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H_2O_2	34.02 g/mol	HO	17.01 g/mol	$n = 2$
C_6H_6	78.11 g/mol	CH	13.02 g/mol	$n = 6$
C_4H_{10}	58.12 g/mol	C_2H_5	29.06 g/mol	$n = 2$
CCl_4	153.81 g/mol	CCl_4	153.81 g/mol	$n = 1$
$C_2H_4Cl_4O_2$	201.85 g/mol	CH_2Cl_2O	100.93 g/mol	$n = 2$