

MODULE 8

The numbers 4 and 12 have a common factor of 4, we divide both subscripts by 4 and get CH_3 . This represents the simplest, whole-number ratio of C atoms and H atoms in the molecule.

The numbers 2, 2, and 8 have a common factor of 2, so we must divide all subscripts by 2 and get NiSO_4 .

The numbers 1, 2, and 7 have no common factors, so $\text{K}_2\text{Cr}_2\text{O}_7$ is already an empirical formula. This particular example shows that empirical formulas can be molecular formulas.

Experiments such as the one in Example 8.9 give only empirical formulas, not molecular formulas. Even though empirical formulas do not contain as much detailed information about a molecule as molecular formulas, they are better at nothing. It is possible to convert a molecule's empirical formula to its molecular formula, if we happen to know the molecule's mass. Think about it: To go from a molecular formula to an empirical formula, what do we have to do? We have to find the common factor and divide by it. For example, in Example 8.9a, we noted that the common factor in the molecular formula was 4; when we divided by 4, we turned the molecular formula into an empirical formula. If we want to turn an empirical formula into a molecular formula, what will we have to do? We will have to multiply by the common factor. How do we know what that common factor is? We can determine it from the molecule's mass. Example 8.10 shows us how to do that.

ON YOUR OWN

8.11 Which of the following molecular formulas are also empirical formulas? If the formula is not an empirical formula, write the empirical formula for that molecule.



EXAMPLE 8.10

A compound's empirical formula is determined by experiment to be CH_2O . In a separate experiment, the molecule's mass was determined to be 180.1 amu. What is the compound's molecular formula?

Remember, to go from an empirical formula to a molecular formula, we must multiply by the common factor. So this problem really boils down to figuring out what the common factor is. We figure that out by looking at the mass. If the molecular formula really were CH_2O , the mass would be:

$$\text{mass of } \text{CH}_2\text{O} = 1 \times 12.0 \text{ amu} + 2 \times 1.01 \text{ amu} + 1 \times 16.0 \text{ amu} = 30.0 \text{ amu}$$

That's anywhere close to the correct mass of 180.1 amu. We need to multiply the numbers in

$2\text{C}_2\text{H}_2 + \text{O}_2 \rightarrow 2\text{C}_2\text{H}_2\text{O} + 2\text{H}_2\text{O}$
If 500.0 grams of $\text{C}_2\text{H}_2\text{O}$ are reacted with excess C_2H_2 , how much H_2O will be produced?

The following reaction is exothermic. How much heat is released when 10.0 g of H_2 reacts with 10.0 g of O_2 ?



1. H_2 is in excess. What is the minimum mass of H_2O that can be produced?



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