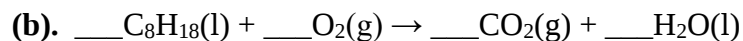
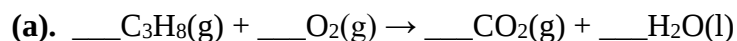


Chapter 6

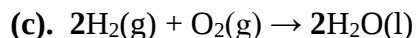
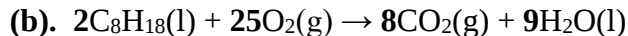
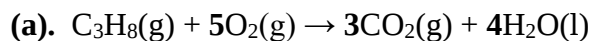
Problems

Problem 6.1

Balance the following equations for the following combustion reactions. Propane (C_3H_8) is used for barbeques and portable stoves; octane (C_8H_{18}) is the major component of gasoline; hydrogen is widely regarded as the fuel of the future.

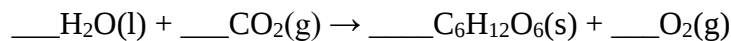


Answer

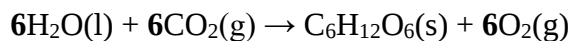


Problem 6.2

Write the balanced equation for photosynthesis, in which water and carbon dioxide are converted to glucose ($\text{C}_6\text{H}_{12}\text{O}_6$) and molecular oxygen (O_2).



Answer



Problem 6.3

Calculate the percentage by mass of carbon and hydrogen in each of the following compounds:

- (a). CH₄
- (b). C₃H₈
- (c). C₈H₁₈
- (d). Glucose

Answer

First obtain the molecular mass. Then apply the equation below.

$$\% \text{ Element} = \frac{(\text{Atoms of element})(\text{Average atomic mass})}{(\text{Molecular mass})} \times 100$$

Note: We will keep an extra digit in the molecular mass, and round to two decimal places for percentages. As expected, the percentages add up to 100.00 for each molecule.

- (a). Methane (CH₄)

$$\begin{aligned} \text{Molecular mass} &= 1 \times (\text{Atomic mass of C}) + 4 \times (\text{Atomic mass of H}) \\ &= 1 \times 12.01 \text{ amu} + 4 \times 1.008 \text{ amu} \\ &= 16.042 \text{ amu} \end{aligned}$$

$$\% \text{ C} = \frac{(1)(12.01 \text{ amu})}{(16.042 \text{ amu})} \times 100 = 74.87\%$$

$$\% \text{ H} = \frac{(4)(1.008 \text{ amu})}{(16.042 \text{ amu})} \times 100 = 25.13\%$$

- (b). Propane (C₃H₈)

$$\begin{aligned} \text{Molecular mass} &= 3 \times (\text{Atomic mass of C}) + 8 \times (\text{Atomic mass of H}) \\ &= 3 \times 12.01 \text{ amu} + 8 \times 1.008 \text{ amu} \end{aligned}$$

$$= 44.094 \text{ amu}$$

$$\% C = \frac{(3)(12.01 \text{ amu})}{(44.094 \text{ amu})} \times 100 = 81.71\%$$

$$\% H = \frac{(8)(1.008 \text{ amu})}{(44.094 \text{ amu})} \times 100 = 18.29\%$$

(c). Octane (C_8H_{18})

$$\text{Molecular mass} = 8 \times (\text{Atomic mass of C}) + 18 \times (\text{Atomic mass of H})$$

$$= 8 \times 12.01 \text{ amu} + 18 \times 1.008 \text{ amu}$$

$$= 114.224 \text{ amu}$$

$$\% C = \frac{(8)(12.01 \text{ amu})}{(114.224 \text{ amu})} \times 100 = 84.12\%$$

$$\% H = \frac{(18)(1.008 \text{ amu})}{(114.224 \text{ amu})} \times 100 = 15.88\%$$

Problem 6.4

In the following figures, atoms are represented by spheres (whole, or partial) with the color codes black for carbon, red for oxygen and white for hydrogen. Calculate the percentage by mass of oxygen in each of the following compounds:

- (a). Carbon monoxide a toxic gas produced upon combustion of fossil fuels when the supply of oxygen is limited.

**Answer**

The molecular formula is CO.

$$\begin{aligned}\text{Molecular mass} &= 1 \times (\text{Atomic mass of C}) + 1 \times (\text{Atomic mass of O}) \\ &= 1 \times 12.01 \text{ amu} + 1 \times 16.00 \text{ amu} \\ &= 28.01 \text{ amu}\end{aligned}$$

$$\% \text{ O} = \frac{(1)(16.00 \text{ amu})}{(28.01 \text{ amu})} \times 100 = 57.12\%$$



- (b).

Formic acid is named after ants which have high concentrations of the compound in their venom. (Wikipedia)

Answer

The molecular formula is H₂CO₂.

$$\begin{aligned}\text{FW} &= 2 \times (\text{Atomic mass of H}) + 1 \times (\text{Atomic mass of C}) + 2 \times (\text{Atomic mass of O}) \\ &= 2 \times 1.008 \text{ amu} + 1 \times 12.01 \text{ amu} + 2 \times 16.00 \text{ amu} \\ &= 46.026 \text{ amu}\end{aligned}$$

$$\% \text{ O} = \frac{(2)(16.00 \text{ amu})}{(46.026 \text{ amu})} \times 100 = 69.53\%$$

Problem 6.5

(a). Calcium carbonate, found in seashells and eggshells. Calcium carbonate is a dietary supplement used when the amount of calcium taken in the diet is not enough. Calcium is needed by the body for healthy bones, muscles, nervous system, and heart. Calcium carbonate is also used as an antacid to relieve heartburn, acid indigestion, and upset stomach. It is available with or without a prescription. (Medline Plus)

Calculate the **molar mass** of calcium carbonate

Answer

From earlier chapters, we recall that the formula for calcium carbonate is CaCO_3 .

$$\begin{aligned} \text{Formula mass} &= 1 \times (\text{Atomic mass of Ca}) + 1 \times (\text{Atomic mass of C}) + 3 \times (\text{Atomic mass of O}) \\ &= 1 \times 40.078 \text{ amu} + 1 \times 12.01 \text{ amu} + 3 \times 16.00 \text{ amu} \\ &= 100.088 \text{ amu, round to } 100.09 \text{ amu} \end{aligned}$$

$$\text{Molar mass} = 100.09 \text{ g/mol.}$$

(b). Calculate the **mass** of 25.0 millimoles (mmol) of calcium carbonate. (1 regular tablet of Tums provides 5.0 mmol of calcium carbonate.)

Answer

Convert mmol to mol and then to mass of CaCO_3 .

$$\begin{aligned} \text{mass of } \text{CaCO}_3 &= 25.0 \text{ mmol } \text{CaCO}_3 \left(\frac{1 \text{ mol } \text{CaCO}_3}{1000 \text{ mmol } \text{CaCO}_3} \right) \left(\frac{100.088 \text{ g } \text{CaCO}_3}{1 \text{ mol } \text{CaCO}_3} \right) \\ &= 2.50 \text{ g } \text{CaCO}_3 \end{aligned}$$

(c). How many oxygen atoms are present in 25.0 millimoles of calcium carbonate?

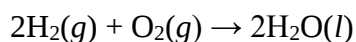
Answer

Convert mmol to mol for CaCO_3 , and then note that 1 mol of CaCO_3 has 3 mol of O

$$\begin{aligned} O \text{ atoms} &= 25.0 \text{ mmol } \text{CaCO}_3 \left(\frac{1 \text{ mol } \text{CaCO}_3}{1000 \text{ mmol } \text{CaCO}_3} \right) \left(\frac{3 \text{ mol } O \text{ atoms}}{1 \text{ mol } \text{CaCO}_3} \right) \left(\frac{6.023 \times 10^{23} O \text{ atoms}}{1 \text{ mol } O \text{ atoms}} \right) \\ &= 4.52 \times 10^{23} O \text{ atoms} \end{aligned}$$

Problem 6.6

Hydrogen is widely regarded as a “fuel of the future”. For the reaction given below, calculate the mass in g of hydrogen that will react completely with 32.0 g of oxygen. What mass of water is produced?



Answer

First, convert mass of O_2 to moles of O_2 , and then use conversion factors based on the balanced equation.

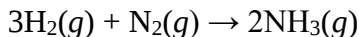
$$\text{mass of } \text{H}_2 = 32.0 \text{ g } \text{O}_2 \left(\frac{1 \text{ mol } \text{O}_2}{32.00 \text{ g } \text{O}_2} \right) \left(\frac{2 \text{ mol } \text{H}_2}{1 \text{ mol } \text{O}_2} \right) \left(\frac{2.016 \text{ g } \text{H}_2}{1 \text{ mol } \text{H}_2} \right) = 4.03 \text{ g } \text{H}_2 \text{ (3sf)}$$

$$\text{mass of } \text{H}_2\text{O} = 32.0 \text{ g } \text{O}_2 \left(\frac{1 \text{ mol } \text{O}_2}{32.00 \text{ g } \text{O}_2} \right) \left(\frac{2 \text{ mol } \text{H}_2\text{O}}{1 \text{ mol } \text{O}_2} \right) \left(\frac{18.016 \text{ g } \text{H}_2\text{O}}{1 \text{ mol } \text{H}_2\text{O}} \right) = 36.0 \text{ g } \text{H}_2\text{O} \text{ (3sf)}$$

Please check: The masses of H_2 and O_2 add up to the mass of H_2O , illustrating the law of conservation of mass!

Problem 6.7

The production of ammonia (NH₃) from nitrogen and hydrogen, shown below, was an important breakthrough resulting in Nobel prizes in chemistry to Fritz Haber (1918) and Carl Bosch (1931). This compound is widely used in manufacturing fertilizers.



Calculate the masses of hydrogen and nitrogen (in kg) needed to manufacture 1000 kg of ammonia. Assume four significant figures.

Answer

It is convenient to work with g, so convert 1000 kg of NH₃ to g first. Then, convert mass of NH₃ to moles of NH₃, and use conversion factors based on the balanced equation. For clarity, convert answers to kg.

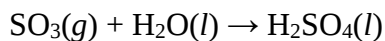
$$\begin{aligned} \text{mass of } H_2 &= 1000 \text{ kg } NH_3 \left(\frac{1000 \text{ g } NH_3}{1 \text{ kg } NH_3} \right) \left(\frac{1 \text{ mol } NH_3}{17.031 \text{ g } NH_3} \right) \left(\frac{3 \text{ mol } H_2}{2 \text{ mol } NH_3} \right) \left(\frac{2.016 \text{ g } H_2}{1 \text{ mol } H_2} \right) \left(\frac{1 \text{ kg } H_2}{1000 \text{ g } H_2} \right) \\ &= 177.6 \text{ kg } H_2 \text{ (4 sf)} \end{aligned}$$

$$\begin{aligned} \text{mass of } N_2 &= 1000 \text{ kg } NH_3 \left(\frac{1000 \text{ g } NH_3}{1 \text{ kg } NH_3} \right) \left(\frac{1 \text{ mol } NH_3}{17.031 \text{ g } NH_3} \right) \left(\frac{1 \text{ mol } N_2}{2 \text{ mol } NH_3} \right) \left(\frac{28.014 \text{ g } N_2}{1 \text{ mol } N_2} \right) \left(\frac{1 \text{ kg } N_2}{1000 \text{ g } N_2} \right) \\ &= 822.4 \text{ kg } N_2 \text{ (4 sf)} \end{aligned}$$

Please check: The two numbers add up to 1000 kg, illustrating the law of conservation of mass.

Problem 6.8

Sulfuric acid (H_2SO_4) is one of the most widely used chemicals in industrial processes.



Calculate the mass of sulfuric acid produced upon the complete reaction of 100.0 kg of sulfur trioxide.

Solution

It is convenient to work with g, so convert 100 kg of SO_3 to g first. Then, convert mass of SO_3 to moles of SO_3 , and use conversion factors based on the balanced equation. For clarity, convert the answer to kg.

$$\begin{aligned} \text{mass of } \text{H}_2\text{SO}_4 &= 100.0 \text{ kg } \text{SO}_3 \left(\frac{1000 \text{ g } \text{SO}_3}{1 \text{ kg } \text{SO}_3} \right) \left(\frac{1 \text{ mol } \text{SO}_3}{80.065 \text{ g } \text{SO}_3} \right) \left(\frac{1 \text{ mol } \text{H}_2\text{SO}_4}{1 \text{ mol } \text{SO}_3} \right) \left(\frac{98.081 \text{ g } \text{H}_2\text{SO}_4}{1 \text{ mol } \text{H}_2\text{SO}_4} \right) \\ &\quad \left(\frac{1 \text{ kg } \text{H}_2\text{SO}_4}{1000 \text{ g } \text{H}_2\text{SO}_4} \right) \\ &= 122.5 \text{ kg } \text{H}_2\text{SO}_4 \text{ (4 sf)} \end{aligned}$$

Problem 6.8

Determine the **empirical formula** of a compound used a fuel in welding with the following composition by mass: 92.26 percent C and 7.74 percent H.

Answer

Assume 100.00 g of the compound, which gives 92.26 g C and 7.74 g H, and calculate moles of each element, and then take ratios.

$$\text{mol of C} = 92.26 \text{ g C} \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 7.68 \text{ mol C}$$

$$\text{mol of H} = 7.74 \text{ g H} \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 7.68 \text{ mol H}$$

The simplest whole number ratios are:

$$\text{C} : = \frac{7.68}{7.68} = 1.00 \qquad \text{H} : = \frac{7.68}{7.68} = 1.00$$

Thus, the empirical formula is C_1H_1 or simply CH, with the subscripts 1 implied.

Problem 6.9

Determine the **empirical formula** of the compound with the following composition by mass: 62.1 percent C, 5.21 percent H, 12.1 percent N and 20.7 percent O.

Answer

Assume 100.00 g of the compound, which gives 62.1 g C and 5.21 g H, 12.1 g N, and 20.7 g O, and calculate moles of each element, and then take ratios.

$$\text{mol of C} = 62.1 \text{ g C} \left(\frac{1 \text{ mol C}}{12.01 \text{ g C}} \right) = 5.17 \text{ mol C}$$

$$\text{mol of H} = 5.21 \text{ g H} \left(\frac{1 \text{ mol H}}{1.008 \text{ g H}} \right) = 5.17 \text{ mol H}$$

$$\text{mol of N} = 12.1 \text{ g N} \left(\frac{1 \text{ mol N}}{14.007 \text{ g N}} \right) = 0.864 \text{ mol N}$$

$$\text{mol of O} = 20.7 \text{ g O} \left(\frac{1 \text{ mol O}}{16.00 \text{ g O}} \right) = 1.29 \text{ mol O}$$

The simplest whole number ratios are obtained by dividing by 0.864, the smallest number:

$$\text{C} := \frac{5.17}{0.864} = 5.98 \quad \text{H} := \frac{5.17}{0.864} = 5.98 \quad \text{N} := \frac{0.864}{0.864} = 1.00 \quad \text{O} := \frac{1.29}{0.864} = 1.49$$

Rounding to the nearest integers, the ratios are 6 for C and H. For O, the number is close to 1.5. Thus, the empirical formula is $\text{C}_6\text{H}_6\text{N}_1\text{O}_{1.5}$. To obtain integers, multiply each subscript by 2 to obtain the empirical formula, $\text{C}_{12}\text{H}_{12}\text{N}_2\text{O}_3$.

Problem 6.10

The empirical formula of a series of compounds is CH_2 . Find the **molecular formulas**:

Molar mass (g/mol)	Molecular Formula
28.052	
84.156	
210.39	

Answer

Let us take the first compound with molar mass 28.052 g/mol, and thus a molecular mass of 28.052 amu

$$\begin{aligned}\text{Empirical formula mass for } \text{CH}_2 &= 1x(\text{Atomic mass of C}) + 2x(\text{Atomic mass of H}) \\ &= 1x12.01 \text{ amu} + 2x1.008 \text{ amu} = 14.026 \text{ amu}\end{aligned}$$

$$\frac{\text{Molecular mass}}{\text{Empirical formula mass}} = \frac{28.052 \text{ amu}}{14.026 \text{ amu}} = 2 \frac{\text{Formula Units}}{\text{Molecule}}$$

The molecular formula is obtained by multiplying each subscript in the empirical formula by 2, noting that the subscript, 1, is implied for carbon in CH_2 . Thus the molecular formula is C_2H_4 .

Verify: Determine the molecular mass for C_2H_4 , and make sure it is 28.052 amu.

$$\begin{aligned}\text{Molecular mass for } \text{C}_2\text{H}_4 &= 2x(\text{Atomic mass of C}) + 4x(\text{Atomic mass of H}) \\ &= 2x12.01 \text{ amu} + 4x1.008 \text{ amu} = 28.052 \text{ amu}\end{aligned}$$

The answers are summarized below:

Molar mass (g/mol)	Molecular mass (amu)	Whole number multiple	Molecular formula
28.052	28.052	2	C ₂ H ₄
84.156	84.156	6	C ₆ H ₁₂
210.39	210.39	15	C ₁₅ H ₃₀

Problem 6.11

As mentioned on the web site of National Propane Association,

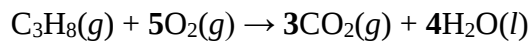
*Propane is more than just outdoor grills. Within the commercial, industrial, and agricultural sectors are many markets in which propane gas is being used. In fact, there are **hundreds of uses of propane**.....*

Calculate the amount of CO₂ produced when 1.00 mole of propane (C₃H₈) is burned in excess of oxygen.

Hint: Begin by writing a **balanced** equation.

Answer

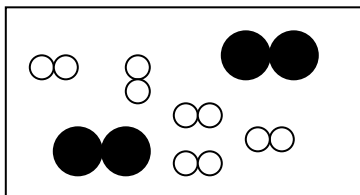
The balanced equation is:



$$\text{mass of CO}_2 = 1.00 \text{ mol C}_3\text{H}_8 \left(\frac{3 \text{ mol CO}_2}{1 \text{ mol C}_3\text{H}_8} \right) \left(\frac{44.01 \text{ g CO}_2}{1 \text{ mol CO}_2} \right) = 132 \text{ g CO}_2 \quad (3\text{sf})$$

Problem 6.12

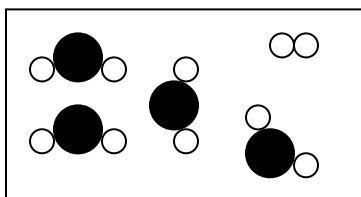
Hydrogen and oxygen react to form water. Consider the mixture of hydrogen and oxygen shown below. The filled circles represent oxygen and the open circles represent hydrogen.



(a). Draw a representation of the product mixture, assuming that the reaction goes to completion.

(b). Identify the reactant that is completely consumed. .

Answer



A water molecule is made of 2 hydrogen atoms and one oxygen atom. The two oxygen molecules (i.e., the four oxygen atoms are used up). Thus oxygen is completely used up. One molecule of hydrogen is in excess.

Problem 6.13

When calcium carbonate is heated vigorously, calcium oxide and carbon dioxide are formed.



When 10.0 g of calcium carbonate is heated, 4.0 g of carbon dioxide is obtained. **Calculate the percent yield.**

Answer

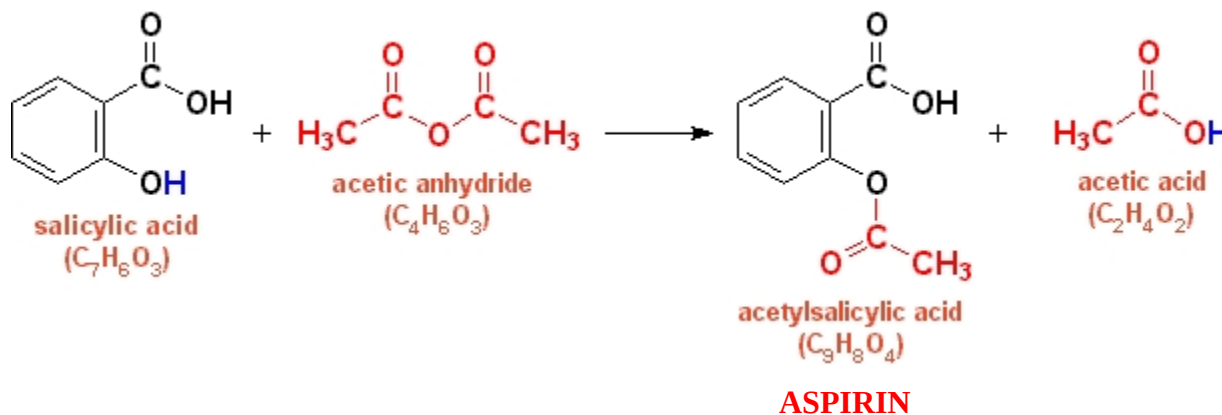
First, calculate the theoretical yield of CO_2 . (**Note:** the molar mass of calcium carbonate is 100.088 g/mol)

$$\text{Theoretical yield of } \text{CO}_2 = 10.0 \text{ g } \text{CaCO}_3 \left(\frac{1 \text{ mol } \text{CaCO}_3}{100.088 \text{ g } \text{CaCO}_3} \right) \left(\frac{1 \text{ mol } \text{CO}_2}{1 \text{ mol } \text{CaCO}_3} \right) \left(\frac{44.01 \text{ g } \text{CO}_2}{1 \text{ mol } \text{CO}_2} \right) = 4.40 \text{ g } \text{CO}_2$$

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{4.0 \text{ g}}{4.40 \text{ g}} \times 100 = 91\% \quad (2\text{sf})$$

Problem 6.14

In an experiment, 2.50 g salicylic acid was reacted with **excess** acetic anhydride to synthesize aspirin, as shown below. In this reaction, 3.02 g of aspirin was obtained. **Calculate the percent yield.**

**Answer**

First, calculate the theoretical yield of aspirin. (**Note:** the molar masses are 138.118 g/mol for $\text{C}_7\text{H}_6\text{O}_3$ and 180.146 g/mol for aspirin.)

$$\begin{aligned}\text{Theoretical yield of aspirin} &= 2.50 \text{ g } \text{C}_7\text{H}_6\text{O}_3 \left(\frac{1 \text{ mol } \text{C}_7\text{H}_6\text{O}_3}{138.118 \text{ g } \text{C}_7\text{H}_6\text{O}_3} \right) \left(\frac{1 \text{ mol aspirin}}{1 \text{ mol } \text{C}_7\text{H}_6\text{O}_3} \right) \left(\frac{180.146 \text{ g aspirin}}{1 \text{ mol aspirin}} \right) \\ &= 3.26 \text{ g aspirin}\end{aligned}$$

$$\text{Percent yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100 = \frac{3.02 \text{ g}}{3.26 \text{ g}} \times 100 = 92.6\% \quad (3 \text{ sf})$$

Problem 6.15

Ingestion of large doses of nitrate (NO_3^-) either in the form of pure sodium nitrate or beetroot juice in young healthy individuals rapidly increases plasma nitrate concentration about 2-3 fold, and this elevated nitrate concentration can be maintained for at least 2 weeks. Increased plasma nitrate stimulates the production of nitric oxide. Nitric oxide (nitrogen monoxide, NO) is important physiological signaling molecule that is used in, among other things, regulation of muscle blood flow and mitochondrial respiration. (Wikipedia)

- (a). The oxidation number of nitrogen in nitrogen monoxide (NO) is _____.
- (b). The oxidation number of nitrogen in the nitrate ion (NO_3^-) (a messenger molecule) is _____.

Answer

a) Oxygen has an oxidation number of -2 . Let us assume that x is the oxidation number of nitrogen. The sum of oxidation numbers is zero for a neutral molecule. Thus,

$$x + (-2) = 0$$

$$x = +2$$

Thus, the oxidation number of nitrogen is $+2$.

b) Oxygen has an oxidation number of -2 . Let us assume that y is the oxidation number of nitrogen. The sum of oxidation numbers is -1 , which is the charge on the nitrate ion.

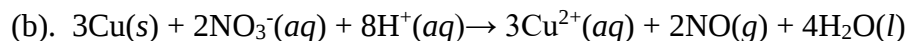
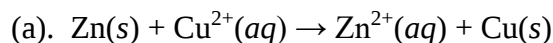
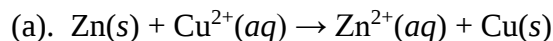
$$y + 3(-2) = -1 \quad (\text{The factor of 3 arises because there are 3 O atoms!})$$

$$y = +5$$

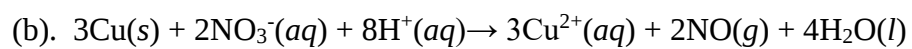
The oxidation number of nitrogen is thus $+5$.

Problem 6.16

True or False. The following are redox reactions

**Answer**

True. We note that the oxidation number of Zn changes from 0 to $+2$ (an increase) and that of Cu changes from $+2$ to 0 (a decrease). Zn is oxidized and Cu is reduced.



True. The oxidation number of Cu changes from 0 to +2, and that for nitrogen changes from +5 to +3, as discussed in Problem 15. Thus, Cu is oxidized and N reduced.