Chapter 3: Mass Relations in Formulas and Chemical Reactions

Section 3.1: The Atomic Mass

The atomic mass is the mass of 1 atom. Atoms are very small and their mass is a very small number. A more practical unit is used to describe the mass of an atom. This unit is called the **atomic mass unit** (expressed as **a.m.u**. or just **u**).

The value of 1 atomic mass unit is chosen as 1/12 of the mass of one carbon-12 isotope. Remember that for carbon-12 the mass number A is equal to 12 (that is carbon-12 has 12 nucleons. The mass of a carbon-12 atom is equal to 1.9926×10^{-23} g. Hence, we conclude 1 u = 1.6605×10^{-24} g. The mass of carbon-12 atom is measured with an instrument called the mass spectrometer.

The atomic masses of elements are generally given in the Periodic Table and are located below the element symbol. For the element carbon, we note that the atomic mass is not 12 but 12.011 u. This is because the element carbon has several isotopes and the number 12.011 u is the average atomic mass of all the isotopes of the element carbon present in a typical sample on earth.

Note: Atomic masses are also called atomic weights.

Section 3.2: The Atomic Mass in the Periodic Table

Check the Periodic Table and look up the atomic mass of different elements by rolling your mouse over the element's symbol.

Section 3.3: Avogadro's Number and the Mole

Typical samples of matter contain huge numbers of atoms, often numbers as large as 10^{24} or more. The **mole** was established as a unit that is very useful when counting the numbers of atoms, ions or molecules. One mole is equal to the number of carbon atoms in 12 g of carbon-12.

1 atom of carbon-12 has a mass of 1.9927×10^{-23} g. Hence, in 12 g of carbon-12, there are:

$$\frac{12\ g}{1.9927 \times 10^{-23}\ g / C\ atom} = 6.022 \times 10^{23}\ C\ atoms$$

The number of 6.022×10^{23} is called **Avogadro's number**. Avogadro's number is expressed by the symbol $\mathbf{N_A}$. Hence, one mole of atoms of carbon-12 (i.e. 12 g of carbon-12) contains Avogadro's number or 6.022×10^{23} atoms of carbon-12.

Note: The term "mole" is analogous to the term "dozen". While a dozen eggs refer to twelve eggs, a mole of particles (atoms, ions or molecules) refers to 6.022×10^{23} particles. It follows that while 2 dozen eggs consists of 24 eggs, 2 moles of particles consists of $2 \times (6.022 \times 10^{23})$ particles (i.e. 1.2044×10^{24} particles).

Sections 3.4 - 3.5: Concept of Formula Mass or Molar Mass

The formula mass, or molar mass, is the sum of atomic masses in a chemical formula.

Examples:

Chemical Formula	Formula Mass (a.m.u.)	Molar Mass (g/mol)
H	1.0	1.0
H_2	2.0	2.0
O	16.0	16.0
O_2	32.0	32.0

Many texts require you to use atomic masses to the nearest hundredth. However, in this DVD you need to always round off to the nearest tenth, unless otherwise specified.

Formula mass is the sum of atomic masses of all atoms in a formula of any molecular or ionic compound. The formula mass is expressed in a.m.u.

Molar mass is the sum of atomic masses of all atoms in a mole of pure substance. The molar mass is expressed in g/mol.

It is important that you know how to write chemical formulas from chemical names and vice versa.

Example: What is the formula mass of tungsten? Click on the Periodic Table. Roll the mouse and find tungsten, W. You will see the mass in red under W. Hence, the formula mass of tungsten, W, is 183.9 a.m.u.

Example: Calculate the molar mass of vitamin A, $C_{20}H_{30}O$.

The vitamin A molecule consists of:

Element	Number of Atoms	Atomic Mass
C	20 x	12.0 = 240.0
H	30 x	1.0 = 30.0
O	1 x	16.0 = 16.0
		286.0

Hence, the molar mass of vitamin A is 286.0 g/mol.

Example: Calculate the molar mass of sodium chloride. Note: Here the chemical name is given but not the chemical formula. Hence, it is important to know the names and formulas of chemicals.

Sodium chloride has the chemical formula NaCl.

Element	Number of Atoms	Atomic Mass
Na	1 x	23.0 = 23.0
Cl	1 x	35.5 = 35.5
		58.5 g/mol

Hence, the molar mass of sodium chloride is 58.5 g/mol.

In Section 3.5, practice Interactive Problems to master these concepts.

Sections 3.6 - 3.8: Conversion between Number of Moles and Mass

When discussing the amount of a substance, it is common practice to use the word “moles” instead of the more rigorous wording “number of moles”. The symbol “n” is often used to describe the number of moles or “moles” of a substance.

How many moles of a substance are present in a given sample of that substance can be calculated from the mass and the molar mass of that substance, according to the equation.

$$\text{Number of moles of a substance is equal to}: n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

Hence, mass = n x molar mass

Thus, if the moles and the chemical formula are given, one can calculate the mass in grams of that chemical.

Example: Calculate the number of moles of NH_3 in 1.0 g of NH_3.

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

The mass of NH_3 is 1.0 g. However, we need to calculate the molar mass of NH_3.

Element	Number of Atoms	Atomic Mass
N	1 x	14.0 = 14.0
H	3 x	1.0 = 3.0
		17.0 g/mol

$$n = \frac{1.0\text{ g}}{17.0\text{ g/mol}} = 0.0588235\text{ mol}$$

In the correct number of significant figures, the answer is 0.059 mol.

Let us continue further with this problem (not covered in the DVD).

Example: How many moles of H and N atoms are present in 1.0 g of NH_3?

From the previous problem, we know that 1.0 g of NH_3 contains 0.059 mol NH_3. From the chemical formula of ammonia, we know that for every molecule of NH_3, there is 1 atom of nitrogen and 3 atoms of hydrogen. So, for every mole of NH_3, there is 1 mole of nitrogen and 3 moles of hydrogen.

Hence, moles of N = moles NH_3 = 0.059 mol N and Moles of H = 3 x moles of NH_3 = 3 x 0.059 mol = 0.18 mol H

Example: Calculate the mass in grams of 12.00 moles of C_2H_3Cl.

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}} \qquad \text{Hence, mass} = n \times \text{molar mass}$$

Calculate the molar mass of C_2H_3Cl.

Element	Number of Atoms	Atomic Mass
C	2 x	12.0 = 24.0
H	3 x	1.0 = 3.0
Cl	1 x	35.5 = 35.5
		62.5 g/mol

mass = 12.00 mol x 62.5 g/mol
mass = 750.g

Example: Calculate the number of moles in 150.0 g of iron(III) oxide.

The chemical formula of iron(III) oxide is Fe_2O_3. In the problem, the mass Fe_2O_3 is given. Hence, the molar mass must be calculated.

Element	Number of Atoms	Atomic Mass
Fe	2 x	55.8 = 111.6
O	3 x	16.0 = 48.0
		159.6 g/mol

$$n = \frac{150.0\text{ g}}{159.6\text{ g/mol}} = 0.93984\text{ mol}$$

In the correct number of significant figures, the answer is 0.9398 mol.

In Section 3.7, visualize one mole for a variety of elements and compounds.

In Section 3.8, practice Interactive Problems.

Sections 3.9 - 3.10: Problems on Avogadro's Number

In some instances, we want to know how many atoms, ions or molecules are involved in a chemical or physical process. We can calculate this number from the mass of the substance, the molar mass of the substance and Avogadro's number.

Avogadro's number is given by the symbol $N_A = 6.022 \times 10^{23}$ or 6.022E23

$$\text{Remember mole, } n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}} \text{ and one mole} = 6.022 \times 10^{23}$$

Example 1: The molecular formula of ethylene glycol is $C_2H_6O_2$. In 13.68 g of ethylene glycol, **(a) calculate the number of moles**

Element	Number of Atoms	Atomic Mass
C	2 x	12.0 = 24.0
H	6 x	1.0 = 6.0
O	2 x	16.0 = 32.0
		62.0

$$n = \frac{13.68 \text{ g}}{62.0 \text{ g/mol}}$$

$$n = 0.221 \text{ mol}$$

Example 1: The molecular formula of ethylene glycol is $C_2H_6O_2$. In 13.68 g of ethylene glycol, **(b) calculate the number of molecules**

$$\frac{6.022 \times 10^{23} \text{ molecules}}{1 \text{mol}} \times 0.221 \text{mol} = 1.33 \times 10^{23} \text{ molecules or } 1.33\text{E}23 \text{ molecules}$$

Example 1: The molecular formula of ethylene glycol is $C_2H_6O_2$. In 13.68 g of ethylene glycol, **(c) Calculate the number of oxygen atoms**

In each molecule of $C_2H_6O_2$, there are 2 atoms of O. Therefore, in 1.33×10^{23} molecules of $C_2H_6O_2$ there are:

$2 \times 1.33 \times 10^{23} = 2.66 \times 10^{23}$ or 2.66E23 atoms of oxygen

Example 2: A sample of the compound, C_3H_6O, contains 14.0×10^{14} carbon atoms. (a) Calculate the number of C_3H_6O molecules

Each molecule of C_3H_6O contains 3 carbon atoms. Therefore, the number of C_3H_6O molecules containing 14.0×10^{14} C atoms is:

$$\frac{14.0 \times 10^{14}}{3} = 4.67 \times 10^{14} \text{ molecules of } C_3H_6O$$

Example 2: A sample of the compound, C_3H_6O, contains 14.0×10^{14} carbon atoms. **(b) Calculate the number of moles of C_3H_6O.**

$$\frac{1 \text{ mol}}{6.022 \times 10^{23} \text{ molecules}} \times 4.67 \times 10^{14} \text{ molecules} = 7.75 \times 10^{-10} \text{ or } 7.75\text{E-}10 \text{ mol}$$

Example2: A sample of the compound, C_3H_6O, contains 14.0×10^{14} carbon atoms. **(c) Calculate the number of grams of C_3H_6O.**

Element	Number of Atoms	Atomic Mass
C	3 x	12.0 = 36.0
H	6 x	1.0 = 6.0
O	1 x	16.0 = 16.0

		58.0 g/mol

The molar mass of C_3H_6O is 58.0 g/mol

$$n = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

Therefore, mass (g) = n x molar mass (g/mol)

Thus, mass = 7.75×10^{-10} mol x 58.0 g/mol
mass = 4.50×10^{-8} or 4.50E-8 g

In Section 3.10, practice the Interactive Problems.

Sections 3.11 - 3.12: Percent Composition

The percent composition of a compound is the mass percent of the elements present.

$$\text{Mass percent of element} = \frac{\text{mass of element in sample}}{\text{total mass of sample}} \times 100$$

Knowing the chemical formula of a compound, the mass percent of its constituent elements can be calculated.

Note: the subscripts in a chemical formula allow us to define the atom ratio as well as the mole ratio in which the different elements are combined.

Example:

In H_2O: the atom ratio is 2 atoms H: 1 atom O
the mole ratio is 2 moles H: 1 mole O

In Na_2SO_4: the atom ratio is 2 atoms Na: 1 atom S: 4 atoms O
the mole ratio is 2 moles Na: 1 mole S: 4 moles O

Example 1: Calculate the mass percent of H and O in H_2O.

Element	Number of Moles	Molar Mass (g/mol)
H	2 x	1.0 = 2.0 g
O	1 x	16.0 = 16.0 g
		18.0 g

$$\text{Mass \% of H} = \frac{2.0\text{ g}}{18.0\text{ g}} \times 100$$
$$= 11\%$$

$$\text{Mass \% O} = \frac{16.0\text{ g}}{18.0\text{ g}} \times 100$$
$$= 88.9\ \%$$

Example 2: Calculate the mass percent of N in C_3H_3N

Element	Number of Moles	Molar Mass (g/mol)
C	3 x	12.0 = 36.0 g
H	3 x	1.0 = 3.0 g
N	1 x	14.0 = 14.0 g
		53.0 g

$$\text{Mass \% N} = \frac{14.0\text{ g}}{53.0\text{ g}} \times 100$$
$$= 26.4\ \%$$

In Section 3.12, practice the Interactive Problems.

Sections 3.13 - 3.14: Empirical Formula

When a new compound is formed or discovered, it is important to determine the chemical formula. Most often, this is done by taking a known amount of sample and decomposing, or breaking down this compound into its component elements.

Or

Taking a known amount of sample and reacting it with oxygen to produce CO_2 and H_2O. The component elements or CO_2 and H_2O, are then collected and weighed. The results of such analyses give the mass of each element in the compound.

This is used to determine the mass percent of each element in the compound. Knowing the mass percent of each element in the compound makes it possible to determine its chemical formula.

Empirical formula is the simplest chemical formula. The simplest formula gives only the ratios of atoms in a compound.

Example: A 50.00 g sample contains 13.28 g of potassium, 17.68 g of chromium, and 19.04 g of oxygen. Find the simplest formula.

Analyze the problem. The sample mass is given, the masses of elements are also given. <u>Make sure</u> the sum of masses of all elements in the sample is equal to the mass of the sample.

Step1: Calculate the number of moles of K, Cr and O in the given masses.

$$\text{Remember that moles} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$13.28\text{ g (K)} \Rightarrow \frac{13.28\text{ g}}{39.1\text{ g/mol}} = 0.340\text{ mol (K)}$$
$$17.68\text{ g (Cr)} \Rightarrow \frac{17.68\text{ g}}{52.0\text{ g/mol}} = 0.340\text{ mol (Cr)}$$
$$19.04\text{ g (O)} \Rightarrow \frac{19.04\text{ g}}{16.0\text{ g/mol}} = 1.19\text{ mol (O)}$$

Step 2: Divide each of the numbers of moles by the smallest number of moles to obtain the relative amounts in moles of each element in the substance.

$$\text{K}: \frac{0.340\text{ mol}}{0.340\text{ mol}} = 1.00$$
$$\text{Cr}: \frac{0.340\text{ mol}}{0.340\text{ mol}} = 1.00$$
$$\text{O}: \frac{1.19\text{ mol}}{0.340\text{ mol}} = 3.50$$

Step 3: Write the formula using these relative numbers of moles of each element. Remember that the subscripts in a formula give the relative numbers of atoms or moles of atoms in that substance.

The results in Step 2 suggest that the simplest formula is: $K_1Cr_1O_{3.5}$

Step 4: Write the final formula, ensuring all subscripts are whole numbers.

We multiply each subscript by 2 to get the empirical formula: $\mathbf{K_2Cr_2O_7}$. This is potassium dichromate. This formula makes sense because the dichromate ion has a -2 charge. The potassium ion has a +1 charge. Hence, this substance has a neutral formula, as it should.

In Section 3.14, practice the Interactive Problem.

Sections 3.15 - 3.16: More Problems on Empirical Formula

To find the composition of a substance, it is often useful to react that substance with oxygen gas. This is an example of **chemical analysis**. Combustion reactions are useful for the chemical analysis of substances containing carbon and hydrogen as they produce CO_2 and H_2O. Measuring the amount of CO_2 and H_2O produced by a given amount of substance allows the determination of how much carbon and hydrogen are present in that amount of substance.

Example: When 5.000 g of ibuprofen is burnt with oxygen gas ($O_{2(g)}$), 13.86 g of $CO_{2(g)}$ and 3.926 g of $H_2O_{(l)}$ are formed. Use the following information to determine the empirical formula of ibuprofen. Ibuprofen is known to contain only carbon, oxygen and hydrogen elements.

Step 1: How much carbon is there in 5.000 g of ibuprofen?
All the carbon in ibuprofen ends up in the 13.86 g of $CO_{2(g)}$. So, the question is how many moles of carbon are present in 13.86 g of $CO_{2(g)}$.

$$13.86\ \text{g}\,(CO_2) \Rightarrow \frac{13.86\ \text{g}\,(CO_2)}{44.01\ \text{g/mol}\ (CO_2)} = 0.3149\ \text{mol}\,(CO_2) \Rightarrow 0.3149\ \text{mol}\,(C)$$

➩ mass (C) in 5 g ibuprofen = 0.3149 mol (C) x 12.01 g/mol (C) = 3.782 g (C)

Step 2: How much hydrogen is there in 5.000 g of ibuprofen?
All the hydrogen in ibuprofen ends up in the 3.926 g of $H_2O_{(l)}$. So, the question is how many moles of hydrogen are present in 3.926 g of $H_2O_{(l)}$.

$$3.926\ \text{g}\,(H_2O) \Rightarrow \frac{3.926\ \text{g}\,(H_2O)}{18.00\ \text{g/mol}\ (H_2O)} = 0.2181\ \text{mol}\,(H_2O) \Rightarrow 0.4362\ \text{mol}\,(H)$$

➩ mass (H) in 5 g ibuprofen = 0.4362 mol (H) x 1.00 g/mol (H) = 0.4362 g (H)

Step 3: How much oxygen is there in 5.000 g of ibuprofen?
Since ibuprofen only contains oxygen, carbon and hydrogen and 5.000 g of ibuprofen contain 3.782 g (C) and 0.4362 g (H), then, the mass of oxygen is:

Mass (O) = 5.000 g – 3.782 g – 0.4362 g = 0.7818 g (O).

Step 4: Now, we can use the strategy shown in Section 3.13 to determine the formula of ibuprofen.

$$C: 3.782\text{ g} \Rightarrow \frac{3.782\text{ g (C)}}{12.0\text{ g/mol (C)}} = 0.3149\text{ mol (C)}$$
$$H: 0.4362\text{ g} \Rightarrow \frac{0.4362\text{ g (H)}}{1.0\text{ g/mol (H)}} = 0.4362\text{ mol (H)}$$
$$O: 0.7818\text{ g} \Rightarrow \frac{0.7818\text{ g (O)}}{16.0\text{ g/mol (O)}} = 0.04886\text{ mol (O)}$$

$$\text{For Carbon} \Rightarrow \frac{0.3149\text{ mol (C)}}{0.04886\text{ mol (O)}} = 6.4\text{ (C) per (O) approximately 6.5}$$
$$\text{For Hydrogen} \Rightarrow \frac{0.4362\text{ mol (H)}}{0.04886\text{ mol (O)}} = 8.9\text{ (H) per (O) approximately 9}$$

Hence, the formula for ibuprofen is: $C_{6.5}H_9O_1$ or more appropriately, $\mathbf{C_{13}H_{18}O_2}$**.**

In Section 3.16, practice the Interactive Problem.

Sections 3.17 - 3.18: Molecular Formula

The empirical formula of a substance is always written using the smallest possible whole number subscripts to give the relative number of atoms of each element in the substance.

Hence, the empirical formula for sodium chloride is written as NaCl and not Na_2Cl_2. NaCl is an ionic compound, not a molecule. Hence, the entity NaCl is called a **formula unit**.

<u>Remember that for ionic compounds, the chemical formula and the empirical formula are always one and the same formula</u>. <u>For molecular (covalent) compounds, however, molecular and empirical formula may be different</u>.

A molecular formula is a whole number multiple of the simplest chemical formula.

Or

A molecular formula is a whole multiple of the empirical formula.

To find the multiple, the molar mass is needed. The empirical formula mass can be calculated from the empirical formula.

$$\text{Multiple} = \frac{\text{molar mass}}{\text{empirical formula mass}}$$

Example: The mass composition of lindane is 24.78% C, 2.08% H and 73.14% Cl. The molar mass of lindane is 290.85 g/mol. Determine the molecular formula.

Element	Mass (g)	Molar Mass (g/mol)	Moles	Mole Ratio
C	24.78	12.0	2.06	1
H	2.08	1.0	2.08	1
Cl	73.14	35.5	2.06	1

Hence, the simplest formula or the empirical formula is CHCl.

The empirical formula mass is 12 + 1 + 35.5 = 48.5 g/mol

$$\text{Multiple} = \frac{290.85\text{ g/mol}}{48.5\text{ g/mol}} = 6$$

Hence, the molecular formula is: $C_6H_6Cl_6$

In Section 3.18, practice the Interactive Problem.

Sections 3.19 - 3.20: Stoichiometry

We are interested in **reaction stoichiometry** whenever we ask questions such as:

1) What is the amount of each reactant required to produce a known amount of product, or,

2) What is the amount of product formed from a known amount of reactants?

Practically speaking, “amounts” of reactants or products are the masses of these compounds measured in the laboratory in grams.

To relate masses of products to masses of reactants requires relating:

1) Masses to moles using the molar masses, and,

2) Moles or reactants to moles of products, using the stoichiometry of the balanced chemical reaction (that is, using the values of the stoichiometric coefficients).

Example: In the Haber process, nitrogen reacts with hydrogen to produce ammonia gas.

$$N_{2(g)} + 3\ H_{2(g)} \longrightarrow 2\ NH_{3(g)}$$

The coefficients (1, 3 and 2) represent the number of moles.

Hence, this equation can be represented as: one mole of N_2 reacts with three moles of H_2 to produce two moles of NH_3.

Whenever we carry out stoichiometric calculations (relating masses of products to masses of reactants), we will always follow the 4 steps given below.

1) Write the balanced chemical equation.

2) Convert masses given for reactants and products to moles, using the molar masses.

3) Write down the **mole ratio** using the stoichiometric coefficients of the balanced chemical equation.

$$\text{Mole Ratio} = \frac{\text{Moles Desired}}{\text{Moles Given}}$$

Moles Desired: Stoichiometric coefficient of the substance for which we want to calculate the amount reacted or produced.

Moles Given: Stoichiometric coefficient of the substance for which we know the amount reacted or produced.

4) Multiply the mole ratio by the number of moles given in the problem or calculated in step 2. Then, convert the calculated number of moles to the mass if necessary.

Example 1: Consider the following reaction: $PH_{3(g)} + O_{2(g)} \rightarrow P_4O_{10(s)} + H_2O_{(l)}$. How many moles of $PH_{3(g)}$ are required for the production of 3.48 mol (P_4O_{10})?

1) Balance the chemical equation:

$$4\ PH_{3(g)} + 8\ O_{2(g)} \rightarrow P_4O_{10(s)} + 6\ H_2O_{(l)}$$

2) Convert masses to moles (already done).

3) Write the Mole Ratio:

“How many moles of $PH_{3(g)}$” implies what is desired is $PH_{3(g)}$.

”Production of 3.48 mol (P_4O_{10})” implies what is given is $P_4O_{10(s)}$.

$$\text{Mole Ratio} = \frac{4\ \text{mol}\,(PH_3)}{1\ \text{mol}\,(P_4O_{10})}$$

4) Multiply the mole ratio by the number of moles given in the problem:

$$\frac{4\ \text{mol}\,(PH_3)}{1\ \text{mol}\,(P_4O_{10})} \times 3.48\ \text{mol}\,(P_4O_{10}) = 13.92\ \text{mol}\ PH_3$$

Example 2: Considering the same chemical reaction, how many moles of $O_{2\,(g)}$ react with 16.3 moles of PH_3?

To solve this problem, let us use the conversion factor (mole ratio).

$$\text{mol}\ O_2 = \frac{8\ \text{mol}\,(O_2)}{4\ \text{mol}\,(PH_3)} \times 16.3\ \text{mol}\ (PH_3) = 32.6\ \text{mol}\ O_2$$

Example 3: Iron reacts with chlorine gas to form iron(III) chloride. Calculate the mass of iron(III) chloride produced from 12.24 moles of iron.

$$2\ Fe_{(s)} + 3\,Cl_{2\,(g)} \longrightarrow 2\ FeCl_{3\,(s)}$$

The reactant Fe is given in moles:

$$\text{In the chemical equation there are:}\ \frac{2\ \text{mol}\ FeCl_3}{2\,\text{mol}\ Fe} \times 12.24\ \text{mol}\ Fe$$
$$= 12.24\ \text{mol}\ FeCl_3$$

However, the problem asks for the mass of $FeCl_3$.

mass = moles x molar mass

mass = 12.24 mol x 162.4 g/mol

mass = 1988 g

Example 4: Silicon dioxide heated with excess of carbon (coke) produces pure silicon and carbon monoxide. How many grams of carbon monoxide are formed when 32.55 g of silicon are produced?

$$SiO_2 + 2\ C \longrightarrow Si + 2\ CO \qquad \frac{32.55\ \text{g Si}}{28.1\ \text{g/mol}} = 1.158\ \text{mol Si}$$

$$\text{In the chemical equation there are: } \frac{2\ \text{mol CO}}{1\ \text{mol Si}} \times 1.158\ \text{mol Si}$$
$$= 2.316\ \text{mol CO}$$

However, the problem asks for the mass of CO.

mass = moles x molar mass
= 2.316 mol x 28.0 g/mol
= 64.90 g CO

In Section 3.20, practice the Interactive Problems.

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