3.1: The Mole

Learning Objectives

- To calculate the molecular mass of a covalent compound and the formula mass of an ionic compound, and to calculate the number of atoms, molecules, or formula units in a sample of a substances.

As discussed previosuly, the **mass number** is the sum of the numbers of protons and neutrons present in the nucleus of an atom. The mass number is an integer that is approximately equal to the numerical value of the atomic mass. Although the mass number is unitless, it is assigned units called **atomic mass units (amu)**. Because a molecule or a polyatomic ion is an assembly of atoms whose identities are given in its molecular or ionic formula, the average atomic mass of any molecule or polyatomic ion can be calculated from its composition by adding together the masses of the constituent atoms. The average mass of a monatomic ion is the same as the average mass of an atom of the element because the mass of electrons is so small that it is insignificant in most calculations.

Molecular and Formula Masses

The molecular mass of a substance is the sum of the average masses of the atoms in one molecule of a substance. It is calculated by adding together the atomic masses of the elements in the substance, each multiplied by its subscript (written or implied) in the molecular formula. Because the units of atomic mass are atomic mass units, the units of molecular mass are also atomic mass units. The procedure for calculating molecular masses is illustrated in Example $3.1.1$.

Example $3.1.1$: Molecular Mass of Ethanol

Calculate the molecular mass of ethanol, whose condensed structural formula is CH_3CH_2OH. Among its many uses, ethanol is a fuel for internal combustion engines.

Given: molecule

Asked for: molecular mass

Strategy:

A. Determine the number of atoms of each element in the molecule.
B. Obtain the atomic masses of each element from the periodic table and multiply the atomic mass of each element by the number of atoms of that element.
C. Add together the masses to give the molecular mass.

Solution:

A The molecular formula of ethanol may be written in three different ways: CH_3CH_2OH (which illustrates the presence of an ethyl group, CH_3CH_2-, and an −OH group), C_2H_5OH, and C_2H_6O; all show that ethanol has two carbon atoms, six hydrogen atoms, and one oxygen atom.

B Taking the atomic masses from the periodic table, we obtain

$$2 \times \text{ atomic mass of carbon } = 2\, atoms \left(\frac{12.011\, amu}{atoms} \right)$$

$$= 24.022\, amu$$

$$6 \times \text{ atomic mass of hydrogen } = 2\, atoms \left(\frac{1.0079\, amu}{atoms} \right)$$

$$= 6.0474\, amu$$

$$1 \times \text{ atomic mass of oxygen } = 1\, atoms \left(\frac{15.9994\, amu}{atoms} \right)$$

$$= 15.994\, amu$$

C Adding together the masses gives the molecular mass:

$$24.022\,amu + 6.0474\,amu + 15.9994\,amu = 46.069\,amu$$

Alternatively, we could have used unit conversions to reach the result in one step:

$$\left[2\,atoms\,C\left(\frac{12.011\,amu}{1\,atom C}\right)\right] + \left[6\,atoms\,H\left(\frac{1.0079\,amu}{1\,atom H}\right)\right] + \left[1\,atoms\,C\left(\frac{15.9994\,amu}{1\,atom 0}\right)\right] = 46.069\,amu$$

The same calculation can also be done in a tabular format, which is especially helpful for more complex molecules:

$$2 \times C \quad (2\,atoms)(12.011\,amu/atom) = 24.022\,amu$$

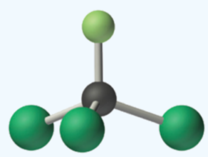
$$6 \times H \quad (6\,atoms)(1.0079\,amu/atom) = 6.0474\,amu$$

$$1 \times O \quad (1\,atoms)(15.9994\,amu/atom) = 15.9994\,amu$$

$$C_2H_6O \quad \text{molecular mass of ethanol} = 46.069\,amu$$

? Exercise 3.1.1: Molecular Mass of Freon

Calculate the molecular mass of trichlorofluoromethane, also known as Freon-11, whose condensed structural formula is $\mathrm{CCl_3F}$. Until recently, it was used as a refrigerant. The structure of a molecule of Freon-11 is as follows:



Freon-11, CCl_3F

Answer

137.368 amu

Unlike molecules, which form covalent bonds, ionic compounds do not have a readily identifiable molecular unit. Therefore, for ionic compounds, the formula mass (also called the empirical formula mass) of the compound is used instead of the molecular mass. The formula mass is the sum of the atomic masses of all the elements in the empirical formula, each multiplied by its subscript (written or implied). It is directly analogous to the molecular mass of a covalent compound. The units are atomic mass units.

> *Atomic mass, molecular mass, and formula mass all have the same units: atomic mass units.*

✓ Example 3.1.2: Formula Mass of Calcium Phosphate

Calculate the formula mass of $\mathrm{Ca_3(PO_4)_2}$, commonly called calcium phosphate. This compound is the principal source of calcium found in bovine milk.

Given: ionic compound

Asked for: formula mass

Strategy:

A. Determine the number of atoms of each element in the empirical formula.
B. Obtain the atomic masses of each element from the periodic table and multiply the atomic mass of each element by the number of atoms of that element.
C. Add together the masses to give the formula mass.

Solution:

A The empirical formula—$Ca_3(PO_4)_2$—indicates that the simplest electrically neutral unit of calcium phosphate contains three Ca^{2+} ions and two PO_4^{3-} ions. The formula mass of this molecular unit is calculated by adding together the atomic masses of three calcium atoms, two phosphorus atoms, and eight oxygen atoms.

B Taking atomic masses from the periodic table, we obtain

$$3 \times \text{atomic mass of calcium} = 3\,atoms\left(\frac{40.078\,amu}{atom}\right) = 120.234\,amu$$

$$2 \times \text{atomic mass of phosphorus} = 2\,atoms\left(\frac{30.973761\,amu}{atom}\right) = 61.947522\,amu$$

$$8 \times \text{atomic mass of oxygen} = 8\,atoms\left(\frac{15.9994\,amu}{atom}\right) = 127.9952\,amu$$

C Adding together the masses gives the formula mass of $\mathrm{Ca_3(PO_4)_2}$:

$$120.234\,amu + 61.947522\,amu + 127.9952\,amu = 310.177\,amu$$

We could also find the formula mass of $\mathrm{Ca_3(PO_4)_2}$ in one step by using unit conversions or a tabular format:

$$\left[3\,atoms Ca\left(\frac{40.078\,amu}{1\,atom Ca}\right)\right] + \left[2\,atoms P\left(\frac{30.973761\,amu}{1\,atom P}\right)\right] + \left[8\,atoms O\left(\frac{15.9994\,amu}{1\,atom O}\right)\right] = 310.177$$

$$amu$$

$$3Ca \quad (3\,atoms)(40.078\,amu/atom) = 120.234\,amu$$

$$2P \quad (2\,atoms)(30.973761\,amu/atom) = 61.947522\,amu$$

$$+8O \quad (8\,atoms)(15.9994\,amu/atom) = 127.9952\,amu$$

$$Ca_3P_2O_8 \quad \text{formula mass of } \mathrm{Ca_3}(PO_4)_2 = 310.177\,amu$$

? Exercise 3.1.2: Formula Mass of Silicon Nitride

Calculate the formula mass of $\mathrm{Si_3N_4}$, commonly called silicon nitride. It is an extremely hard and inert material that is used to make cutting tools for machining hard metal alloys.

Answer

$140.29\,amu$

Molar Masses of Compounds: Molar Masses of Compounds, YouTube(opens in new window) [youtu.be]

The Mole

Dalton's theory that each chemical compound has a particular combination of atoms and that the ratios of the numbers of atoms of the elements present are usually small whole numbers. It also describes the law of multiple proportions, which states that the ratios of the masses of elements that form a series of compounds are small whole numbers. The problem for Dalton and other early chemists was to discover the quantitative relationship between the number of atoms in a chemical substance and its mass. Because the masses of individual atoms are so minuscule (on the order of 10^{-23} g/atom), chemists do not measure the mass of individual atoms or molecules. In the laboratory, for example, the masses of compounds and elements used by chemists typically range from milligrams to grams, while in industry, chemicals are bought and sold in kilograms and tons. To analyze the transformations that occur between individual atoms or molecules in a chemical reaction, it is therefore essential for chemists to know how many atoms or molecules are contained in a measurable quantity in the laboratory—a given mass of sample. The unit that provides this link is the mole (mol), from the Latin *moles*, meaning "pile" or "heap."

Many familiar items are sold in numerical quantities with distinct names. For example, cans of soda come in a six-pack, eggs are sold by the dozen (12), and pencils often come in a gross (12 dozen, or 144). Sheets of printer paper are packaged in reams of 500, a seemingly large number. Atoms are so small, however, that even 500 atoms are too small to see or measure by most common techniques. Any readily measurable mass of an element or compound contains an extraordinarily large number of atoms, molecules, or ions, so an extremely large numerical unit is needed to count them. The mole is used for this purpose.

A mole is defined as the amount of a substance that contains the number of carbon atoms in exactly 12 g of isotopically pure carbon-12. According to the most recent experimental measurements, this mass of carbon-12 contains 6.022142×10^{23} atoms, but for most purposes 6.022×10^{23} provides an adequate number of significant figures. Just as 1 mole of atoms contains 6.022×10^{23} atoms, 1 mole of eggs contains 6.022×10^{23} eggs. This number is called Avogadro's number, after the 19th-century Italian scientist who first proposed a relationship between the volumes of gases and the numbers of particles they contain.

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.022×10^{23}.

> *One mole always has the same number of objects: 6.022 × 1023.*

To appreciate the magnitude of Avogadro's number, consider a mole of pennies. Stacked vertically, a mole of pennies would be 4.5×10^{17} mi high, or almost six times the diameter of the Milky Way galaxy. If a mole of pennies were distributed equally among the entire population on Earth, each person would have more than one trillion dollars. The mole is so large that it is useful only for measuring very small objects, such as atoms.

The concept of the mole allows scientists to count a specific number of individual atoms and molecules by weighing measurable quantities of elements and compounds. To obtain 1 mol of carbon-12 atoms, one weighs out 12 g of isotopically pure carbon-12. Because each element has a different atomic mass, however, a mole of each element has a different mass, even though it contains the same number of atoms (6.022×10^{23}). This is analogous to the fact that a dozen extra large eggs weighs more than a dozen small eggs, or that the total weight of 50 adult humans is greater than the total weight of 50 children. Because of the way the mole is defined, for every element the number of grams in a mole is the same as the number of atomic mass units in the atomic mass of the element. For example, the mass of 1 mol of magnesium (atomic mass = 24.305 amu) is 24.305 g. Because the atomic mass of magnesium (24.305 amu) is slightly more than twice that of a carbon-12 atom (12 amu), the mass of 1 mol of magnesium atoms (24.305 g) is slightly more than twice that of 1 mol of carbon-12 (12 g). Similarly, the mass of 1 mol of helium (atomic mass = 4.002602 amu) is 4.002602 g, which is about one-third that of 1 mol of carbon-12. Using the concept of the mole, Dalton's theory can be restated: 1 mol of a compound is formed by combining elements in amounts whose mole ratios are small whole numbers. For example, 1 mol of water (H_2O) has 2 mol of hydrogen atoms and 1 mol of oxygen atoms.

Molar Mass

The molar mass of a substance is defined as the mass in grams of 1 mole of that substance. One mole of isotopically pure carbon-12 has a mass of 12 g. For an element, the molar mass is the mass of 1 mol of atoms of that element; for a covalent molecular compound, it is the mass of 1 mol of molecules of that compound; for an ionic compound, it is the mass of 1 mol of formula units. That is, the molar mass of a substance is the mass (in grams per mole) of 6.022×10^{23} atoms, molecules, or formula units of that substance. In each case, the number of grams in 1 mol is the same as the number of atomic mass units that describe the atomic mass, the molecular mass, or the formula mass, respectively.

> *The molar mass of any substance is its atomic mass, molecular mass, or formula mass in grams per mole.*

The periodic table lists the atomic mass of carbon as 12.011 amu; the average molar mass of carbon—the mass of 6.022 × 10^{23} carbon atoms—is therefore 12.011 g/mol:

Atomic mass of carbon as 12.011 amu; the average molar mass of carbon

Substance (formula)	Atomic, Molecular, or Formula Mass (amu)	Molar Mass (g/mol)
carbon (C)	12.011 (atomic mass)	12.011
ethanol (C_2H_5OH)	46.069 (molecular mass)	46.069
calcium phosphate [$Ca_3(PO_4)_2$]	310.177 (formula mass)	310.177

Determining the Molar Mass of a Molecule, YouTube: Determining the Molar Mass of a Molecule(opens in new window) [youtu.be]

The molar mass of naturally-occurring carbon is different from that of carbon-12, and is not an integer because carbon occurs as a mixture of carbon-12, carbon-13, and carbon-14. One mole of carbon still has 6.022 × 10^{23} carbon atoms, but 98.89% of those atoms are carbon-12, 1.11% are carbon-13, and a trace (about 1 atom in 1012) are carbon-14. (For more information, see Section 1.6.) Similarly, the molar mass of uranium is 238.03 g/mol, and the molar mass of iodine is 126.90 g/mol. When dealing with elements such as iodine and sulfur, which occur as a diatomic molecule (I_2) and a polyatomic molecule (S_8), respectively, molar mass usually refers to the mass of 1 mol of atoms of the element—in this case I and S, not to the mass of 1 mol of molecules of the element (I_2 and S_8).

The molar mass of ethanol is the mass of ethanol (C_2H_5OH) that contains 6.022 × 10^{23} ethanol molecules. As in Example $3.1.1$, the molecular mass of ethanol is 46.069 amu. Because 1 mol of ethanol contains 2 mol of carbon atoms (2 × 12.011 g), 6 mol of hydrogen atoms (6 × 1.0079 g), and 1 mol of oxygen atoms (1 × 15.9994 g), its molar mass is 46.069 g/mol. Similarly, the formula mass of calcium phosphate [$Ca_3(PO_4)_2$] is 310.177 amu, so its molar mass is 310.177 g/mol. This is the mass of calcium phosphate that contains 6.022 × 10^{23} formula units.

The mole is the basis of quantitative chemistry. It provides chemists with a way to convert easily between the mass of a substance and the number of individual atoms, molecules, or formula units of that substance. Conversely, it enables chemists to calculate the mass of a substance needed to obtain a desired number of atoms, molecules, or formula units. For example, to convert moles of a substance to mass, the following relationship is used:

$$(\text{moles}) \times (\text{molar mass}) \rightarrow \text{mass} \tag{3.1.1}$$

or, more specifically,

$$\cancel{\text{moles}} \times \left(\frac{\text{grams}}{\cancel{\text{mole}}}\right) = \text{grams}$$

Conversely, to convert the mass of a substance to moles:

$$\left(\frac{\text{grams}}{\text{molar mass}}\right) \rightarrow \text{moles} \tag{3.1.2}$$

$$\left(\frac{\text{grams}}{\text{grams/mole}}\right) = \cancel{\text{grams}}\left(\frac{\text{mole}}{\cancel{\text{grams}}}\right) = \text{moles} \tag{3.1.3}$$

The coefficients in a balanced chemical equation can be interpreted both as the relative numbers of molecules involved in the reaction and as the relative number of moles. For example, in the *balanced* equation:

$$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(\ell)$$

the production of two moles of water would require the consumption of 2 moles of H_2 and one mole of O_2. Therefore, when considering ***this*** particular reaction

- **2 moles** of H_2
- **1 mole** of O_2 and
- **2 moles** of H_2O

would be considered to be ***stoichiometrically equivalent quantitites***.

These stoichiometric relationships, derived from balanced equations, can be used to determine expected amounts of products given amounts of reactants. For example, how many moles of H_2O would be produced from 1.57 moles of O_2?

$$(1.57\ mol\ O_2)\left(\frac{2\ mol\ H_2O}{1\ mol\ O_2}\right) = 3.14\ mol\ H_2O$$

The ratio $\left(\frac{2\ mol\ H_2O}{1\ mol\ O_2}\right)$ is the stoichiometric relationship between H_2O and O_2 from the balanced equation for this reaction.

Be sure to pay attention to the units when converting between mass and moles. Figure 3.1.1 is a flowchart for converting between mass; the number of moles; and the number of atoms, molecules, or formula units. The use of these conversions is illustrated in Examples 3.1.3 and 3.1.4.

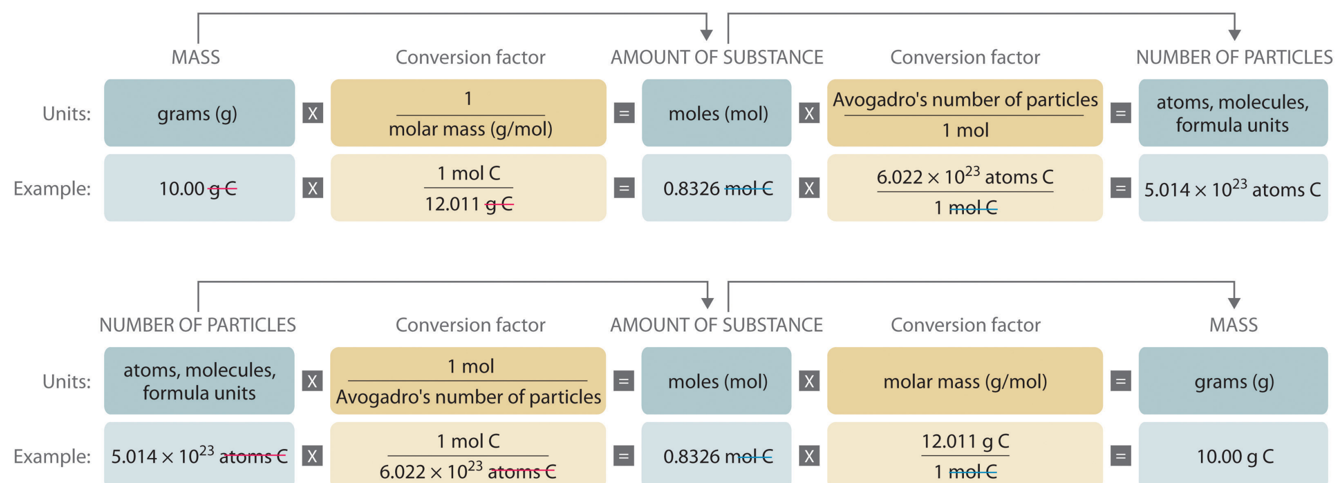


Figure 3.1.1: A Flowchart for Converting between Mass; the Number of Moles; and the Number of Atoms, Molecules, or Formula Units

Flowchart of conversions between mass, amount of substance, and number of particles using conversion factors of molar mass and Avogadro's number.

Example 3.1.3: Combustion of Butane

For the combustion of butane (C_4H_{10}) the balanced equation is:

$$2\,C_4H_{10}(l) + 13\,O_2(g) \rightarrow 8\,CO_2(g) + 10\,H_2O(l)$$

Calculate the mass of CO_2 that is produced in burning 1.00 gram of C_4H_{10}.

Solution

Thus, the overall sequence of steps to solve this problem is:

grams of reactant → moles of reactant → moles of product → grams of product

First of all we need to calculate how many moles of butane we have in a 1.00 gram sample:

$$(1.00\ g\ C_4H_{10})\left(\frac{1\ mol\ C_4H_{10}}{58.0\ g\ C_4H_{10}}\right) = 1.72 \times 10^{-2}\ mol\ C_4H_{10}$$

Now, the stoichiometric relationship between C_4H_{10} and CO_2 is:

$$\left(\frac{8\ mol\ CO_2}{2\ mol\ C_4H_{10}}\right)$$

Therefore:

$$\left(\frac{8\ mol\ CO_2}{2\ mol\ C_4H_{10}}\right) \times 1.72 \times 10^{-2}\ mol\ C_4H_{10} = 6.88 \times 10^{-2}\ mol\ CO_2$$

The question called for the determination of the mass of CO_2 produced, thus we have to convert moles of CO_2 into grams (by using the **molecular weight** of CO_2):

$$6.88 \times 10^{-2}\ mol\ CO_2 \left(\frac{44.0\ g\ CO_2}{1\ mol\ CO_2}\right) = 3.03\ g\ CO_2$$

Example 3.1.4: Ethylene Glycol

For 35.00 g of ethylene glycol (\ce{HOCH2CH2OH}), which is used in inks for ballpoint pens, calculate the number of

a. moles.
b. molecules.

Given: mass and molecular formula

Asked for: number of moles and number of molecules

Strategy:

A. Use the molecular formula of the compound to calculate its molecular mass in grams per mole.
B. Convert from mass to moles by dividing the mass given by the compound's molar mass.
C. Convert from moles to molecules by multiplying the number of moles by Avogadro's number.

Solution:

A The molecular mass of ethylene glycol can be calculated from its molecular formula using the method illustrated in Example 3.1.1:

$$2C(2\,atoms)(12.011\,amu/atom) = 24.022\,amu$$

$$6H(6\,atoms)(1.0079\,amu/atom) = 6.0474\,amu$$

$$2O(2\,atoms)(15.9994\,amu/atom) = 31.9988\,amu$$

$$C_2H_6O_2 \text{molecular mass of ethylene glycol} = 62.068\,amu$$

The molar mass of ethylene glycol is 62.068 g/mol.

B The number of moles of ethylene glycol present in 35.00 g can be calculated by dividing the mass (in grams) by the molar mass (in grams per mole):

$$\frac{\text{mass of ethylene glycol (g)}}{\text{molar mass (g/mol)}} = \text{moles ethylene glycol (mol)}$$

So

$$35.00\, g \text{ethylene glycol} \left(\frac{1\, mole \text{ ethylene glycol}}{62.068\, g \text{ ethylene glycol}} \right) = 0.5639\, mol \text{ethylene glycol}$$

It is always a good idea to estimate the answer before you do the actual calculation. In this case, the mass given (35.00 g) is less than the molar mass, so the answer should be less than 1 mol. The calculated answer (0.5639 mol) is indeed less than 1 mol, so we have probably not made a major error in the calculations.

C To calculate the number of molecules in the sample, we multiply the number of moles by Avogadro's number:

$$\text{molecules of ethylene glycol} = 0.5639\ \cancel{mol} \left(\frac{6.022 \times 10^{23}\, molecules}{1\ \cancel{mol}} \right)$$

$$= 3.396 \times 10^{23}\, molecules$$

Because we are dealing with slightly more than 0.5 mol of ethylene glycol, we expect the number of molecules present to be slightly more than one-half of Avogadro's number, or slightly more than 3×10^{23} molecules, which is indeed the case.

? Exercise 3.1.4: Freon-11

For 75.0 g of CCl_3F (Freon-11), calculate the number of

a. moles.
b. molecules.

Answer a

0.546 mol

Answer b

3.29×10^{23} molecules

✔ Example 3.1.5

Calculate the mass of 1.75 mol of each compound.

a. S_2Cl_2 (common name: sulfur monochloride; systematic name: disulfur dichloride)
b. $Ca(ClO)_2$ (calcium hypochlorite)

Given: number of moles and molecular or empirical formula

Asked for: mass

Strategy:

A Calculate the molecular mass of the compound in grams from its molecular formula (if covalent) or empirical formula (if ionic).

B Convert from moles to mass by multiplying the moles of the compound given by its molar mass.

Solution:

We begin by calculating the molecular mass of S_2Cl_2 and the formula mass of $Ca(ClO)_2$.

A The molar mass of S_2Cl_2 is obtained from its molecular mass as follows:

$$2S(2\,atoms)(32.065\,amu/atom) = 64.130\,amu$$

$$+2Cl(2\,atoms)(35.453\,amu/atom) = 70.906\,amu$$

$$S_2Cl_2\text{molecular mass of }S_2Cl_2 = 135.036\,amu$$

The molar mass of S_2Cl_2 is 135.036 g/mol.

B The mass of 1.75 mol of S_2Cl_2 is calculated as follows:

$$moles S_2Cl_2\left[\text{molar mass}\left(\frac{g}{mol}\right)\right] \rightarrow mass of\, S_2Cl_2\,(g)$$

$$1.75\,mol S_2Cl_2\left(\frac{135.036\,g S_2Cl_2}{1\,mol S_2Cl_2}\right) = 236\,g S_2Cl_2$$

A The formula mass of $Ca(ClO)_2$ is obtained as follows:

$$1Ca(1\,atom)(40.078\,amu/atom) = 40.078\,amu$$

$$2Cl(2\,atoms)(35.453\,amu/atom) = 70.906\,amu$$

$$+2O(2\,atoms)(15.9994\,amu/atom) = 31.9988\,amu$$

$$Ca(ClO)_2\text{ formula mass of }Ca(ClO)_2 = 142.983\,amu$$

The molar mass of $Ca(ClO)_2$ 142.983 g/mol.

B The mass of 1.75 mol of $Ca(ClO)_2$ is calculated as follows:

$$moles Ca(ClO)_2\left[\frac{\text{molar mass}Ca(ClO)_2}{1\,mol Ca(ClO)_2}\right] = mass Ca(ClO)_2$$

$$1.75\,mol Ca(ClO)_2\left[\frac{142.983\,g Ca(ClO)_2}{1\,mol Ca(ClO)_2}\right] = 250\,g Ca(ClO)_2$$

Because 1.75 mol is less than 2 mol, the final quantity in grams in both cases should be less than twice the molar mass, which it is.

Exercise 3.1.5

Calculate the mass of 0.0122 mol of each compound.

a. Si_3N_4 (silicon nitride), used as bearings and rollers
b. $(CH_3)_3N$ (trimethylamine), a corrosion inhibitor

Answer a

1.71 g

Answer b

0.721 g

Conversions Between Grams, Mol, & Atoms: Conversions Between Grams, Mol & Atoms, YouTube(opens in new window) [youtu.be]

The coefficients in a balanced chemical equation can be interpreted both as the relative numbers of molecules involved in the reaction and as the relative number of moles. For example, in the *balanced* equation:

$$2\,\mathrm{H_2(g)} + \mathrm{O_2(g)} \rightarrow 2\,\mathrm{H_2O(l)}$$

the production of two moles of water would require the consumption of 2 moles of H_2 and one mole of O_2. Therefore, when considering ***this*** particular reaction

- **2 moles** of H_2
- **1 mole** of O_2 and
- **2 moles** of H_2O

would be considered to be ***stoichiometrically equivalent quantitites***.

These stoichiometric relationships, derived from balanced equations, can be used to determine expected amounts of products given amounts of reactants. For example, how many moles of H_2O would be produced from 1.57 moles of O_2?

$$(1.57\ mol\ O_2)\left(\frac{2\ mol H_2O}{1\ mol\ O_2}\right) = 3.14\ mol\ H_2O$$

The ratio $\left(\frac{2\ mol\backslash 1 H_2O}{1\ mol\ O_2}\right)$ is the stoichiometric relationship between H_2O and O_2 from the balanced equation for this reaction.

Example 3.1.6

For the combustion of butane (C_4H_{10}) the balanced equation is:

$$2\,\mathrm{C_4H_{10}(l)} + 13\,\mathrm{O_2(g)} \rightarrow 8\,\mathrm{CO_2(g)} + 10\,\mathrm{H_2O(l)}$$

Calculate the mass of CO_2 that is produced in burning 1.00 gram of C_4H_{10}.

Solution

First of all we need to calculate how many moles of butane we have in a 1.00 gram sample:

$$(1.00\ g\ C_4H_{10})\left(\frac{1\ mol\ C_4H_{10}}{58.0\ g\ C_4H_{10}}\right) = 1.72 \times 10^{-2}\ mol\ C_4H_{10}$$

Now, the stoichiometric relationship between C_4H_{10} and CO_2 is:

$$\left(\frac{8\ mol\ CO_2}{2\ mol\ C_4H_{10}}\right)$$

Therefore:

$$\left(\frac{8\ mol\ CO_2}{2\ mol\ C_4H_{10}}\right) \times 1.72 \times 10^{-2}\ mol\ C_4H_{10} = 6.88 \times 10^{-2}\ mol\ CO_2$$

The question called for the determination of the mass of CO_2 produced, thus we have to convert moles of CO_2 into grams (by using the **molecular weight** of CO_2):

$$6.88 \times 10^{-2}\ mol\ CO_2 \left(\frac{44.0\ g\ CO_2}{1\ mol\ CO_2}\right) = 3.03\ g\ CO_2$$

Thus, the overall sequence of steps to solve this problem were:

grams of reactant → moles of reactant → moles of product → grams of product

In a similar way we could determine the mass of water produced, or oxygen consumed, etc.

Finding Mols and Masses of Reactants and Products Using Stoichiometric Factors (Mol Ratios): Finding Mols and Masses of Reactants and Products Using Stoichiometric Factors (Mol Ratios), YouTube(opens in new window) [youtu.be]

Summary

To analyze chemical transformations, it is essential to use a standardized unit of measure called the mole. The molecular mass and the formula mass of a compound are obtained by adding together the atomic masses of the atoms present in the molecular formula or empirical formula, respectively; the units of both are atomic mass units (amu). The mole is a unit used to measure the number of atoms, molecules, or (in the case of ionic compounds) formula units in a given mass of a substance. The mole is defined as the amount of substance that contains the number of carbon atoms in exactly 12 g of carbon-12, Avogadro's number (6.022×10^{23}) of atoms of carbon-12. The molar mass of a substance is defined as the mass of 1 mol of that substance, expressed in grams per mole, and is equal to the mass of 6.022×10^{23} atoms, molecules, or formula units of that substance.

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