Unit 3 includes the following sections of OpenStax

- 3.1 Formula Mass and the Mole Concept
- 3.2 Determining Empirical and Molecular Formulas
- 3.3 Molarity

Figure 3.1

The water in a swimming pool is a complex mixture of substances whose relative amounts must be carefully maintained to ensure the health and comfort of people using the pool. (credit: modification of work by Vic Brincat)

Learning Objectives

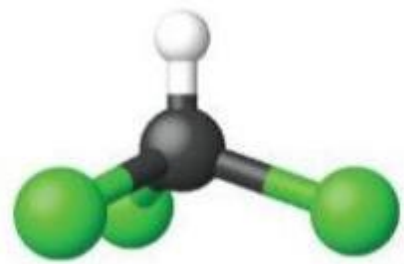
- 3.1 Formula Mass and the Mole Concept
  - Calculate formula masses for covalent and ionic compounds
  - Define the amount unit mole and the related quantity Avogadro's number Explain the relation between mass, moles, and numbers of atoms or molecules, and perform calculations deriving these quantities from one another

Formula Mass and the Mole Concept

- The **formulas mass** of a substance is the sum of the average atomic masses of all the atoms in the substance’s formula.
- Covalent substances exist as discrete molecules.
- The formula mass of a covalent substance may be correctly referred to as a **molecular mass**.

Figure 3.2

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
			Molecular mass		119.37

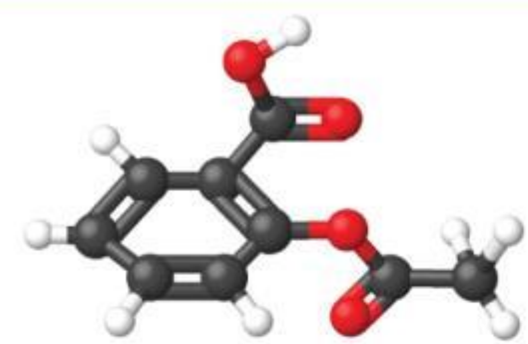


The average mass of a chloroform molecule, $CHCl_3$, is 119.37 amu, which is the sum of the average atomic masses of each of its constituent atoms. The model shows the molecular structure of chloroform.

Figure 3.3

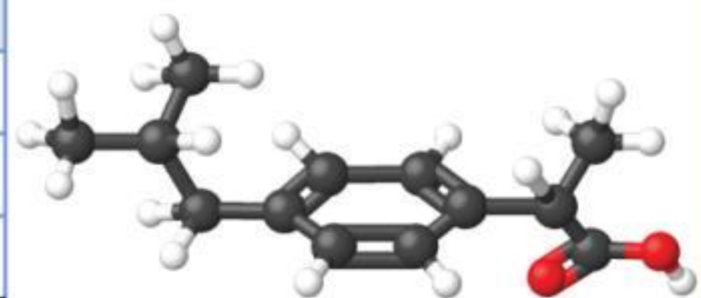
- The average mass of an aspirin molecule is 180.15 amu. The model shows the molecular structure of aspirin, $C_9H_8O_4$.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	9	×	12.01	=	108.09
H	8	×	1.008	=	8.064
O	4	×	16.00	=	64.00
			Molecular mass		180.15



Example 3.1

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	13	×	12.01	=	156.13
H	18	×	1.008	=	18.144
O	2	×	16.00	=	32.00
				Molecular mass	206.27



Formula Mass for Ionic Compounds

- Ionic substances are composed of discrete cations and anions combined in ratios to yield electrically neutral bulk matter.
- Ionic compounds do ***not*** exist as molecules.
- The formula mass for an ionic compound may ***not*** correctly be referred to as a molecular mass.
- The average atomic masses of the ions can be approximated to be the same as the average atomic masses of the neutral atoms.

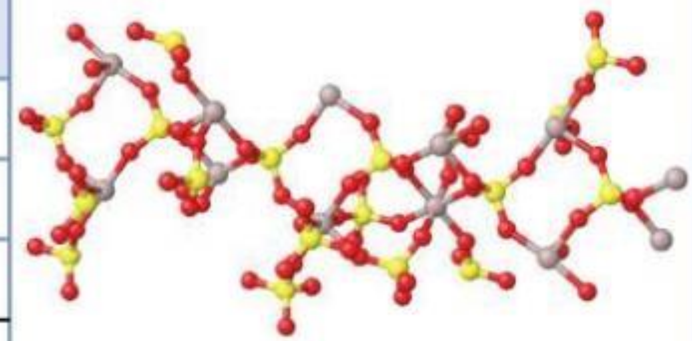
Figure 3.4

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
			Formula mass		58.44

Table salt, NaCl, contains an array of sodium and chloride ions combined in a 1:1 ratio. Its formula mass is 58.44 amu.

Example 3.2

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
Al	2	×	26.98	=	53.96
S	3	×	32.06	=	96.18
O	12	×	16.00	=	192.00
			Molecular mass		342.14



The Mole

- The ***mole*** is an amount unit similar to familiar units like pair, dozen, gross, etc.

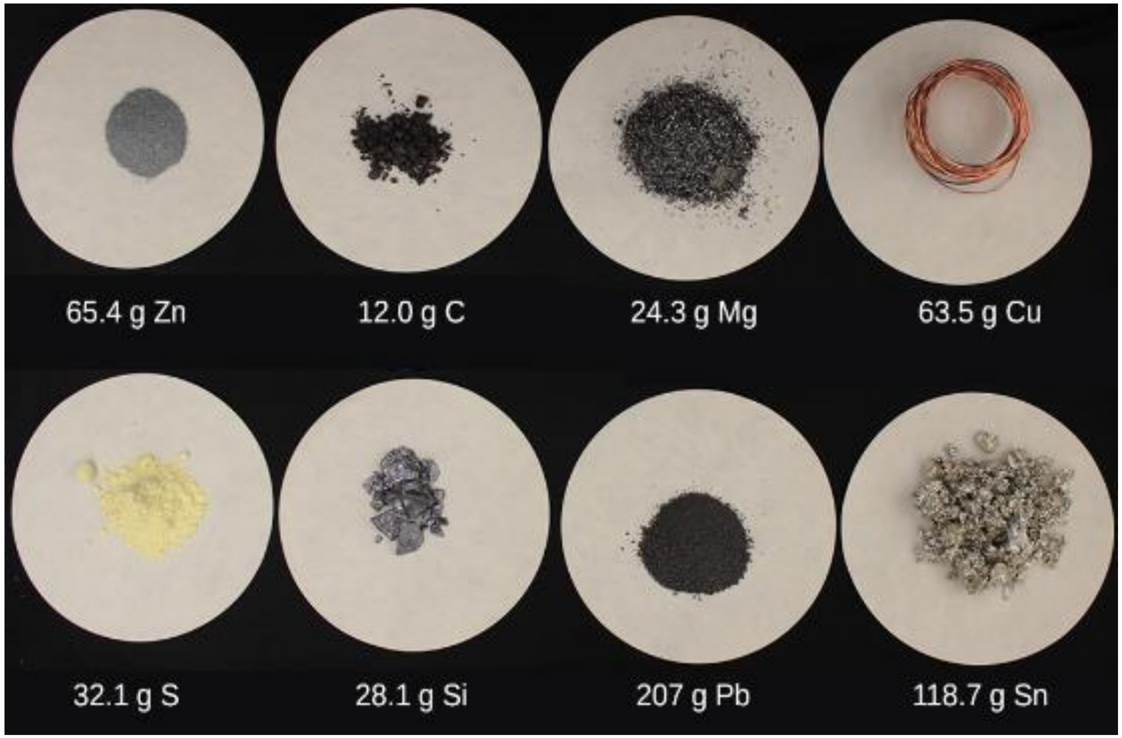
- The mole is defined as the amount of a substance containing the same number of discrete entities (such as atoms, molecules, or ions) as the number of atoms in a sample of pure carbon-12 weighing exactly 12 g.

- The mole provides a link between the mass of a sample and the number of atoms, molecules, or ions in that sample.

Avogadro's Number

- The number of entities composing a mole has been determined to be $6.02214179 \times 10^{23}$.
- This constant is named after Italian scientist Amedeo Avogadro and is known as ***Avogadro's Number***.
- Avogadro's Number (N_A) = 6.022×10^{23}
- The masses of 1 mole of different elements, however, are different, since the masses of the individual atoms are drastically different.

Figure 3.5



Each sample contains 6.022×10^{23} atoms —1.00 mol of atoms. From left to right (top row): 65.4 g zinc, 12.0 g carbon, 24.3 g magnesium, and 63.5 g copper. From left to right (bottom row): 32.1 g sulfur, 28.1 g silicon, 207 g lead, and 118.7 g tin. (credit: modification of work by Mark Ott)

Molar Mass

- The **molar mass** of an element (or compound) is the mass in grams of 1 mole of that substance, a property expressed in units of grams per mole (g/mol).

- The molar mass of any substance is numerically equivalent to its atomic or formula mass in amu.

- Example:
  - A single 12C atom has a mass of 12 amu.
  - A mole of 12C atoms have a mass of 12 g.

Figure 3.6

Each sample contains 6.02×10^{23} molecules or formula units—1.00 mol of the compound or element. Clock-wise from the upper left: 130.2 g of $C_8H_{17}OH$ (1-octanol, formula mass 130.2 amu), 454.9 g of HgI_2 (mercury(II) iodide, formula mass 459.9 amu), 32.0 g of CH_3OH (methanol, formula mass 32.0 amu) and 256.5 g of S_8 (sulfur, formula mass 256.6 amu). (credit: Sahar Atwa)

Molar Mass of Elements

Element	Average Atomic Mass (amu)	Molar Mass (g/mol)	Atoms/Mole
C	12.01	12.01	6.022×10^{23}
H	1.008	1.008	6.022×10^{23}
O	16.00	16.00	6.022×10^{23}
Na	22.99	22.99	6.022×10^{23}
Cl	35.45	35.45	6.022×10^{23}

Figure 3.7

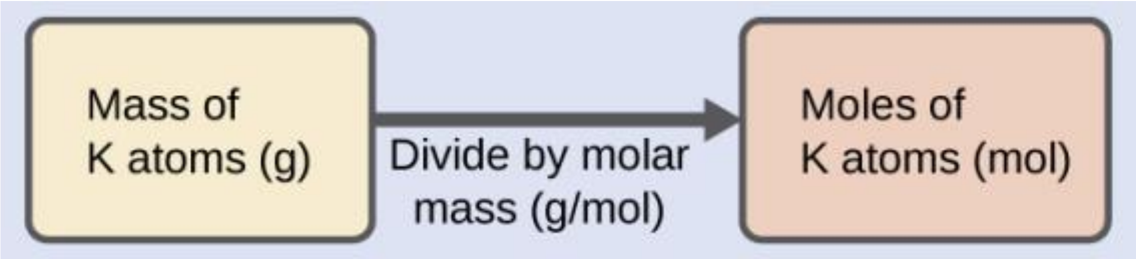
The number of molecules in a single droplet of water is roughly 100 billion times greater than the number of people on earth. (credit: “tanakawho”/Wikimedia commons)

Calculations

- The relationships between formula mass, the mole, and Avogadro's number can be applied to compute various quantities that describe the composition of substances and compounds.

Example 3.3

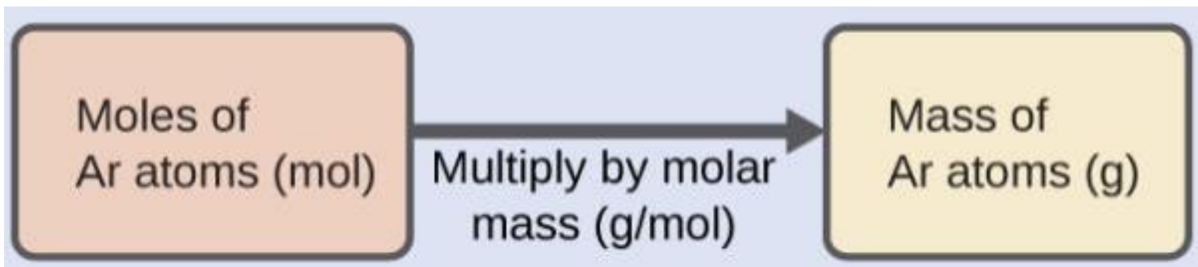
According to nutritional guidelines from the US Department of Agriculture, the estimated average requirement for dietary potassium is 4.7 g. What is the estimated average requirement of potassium in moles?



$$4.7\ \cancel{\text{g}}\ \text{K}\left(\frac{1\ \text{mol K}}{39.10\ \cancel{\text{g}}}\right) = 0.12\ \text{mol K}$$

Example 3.4

A liter of air contains 9.2 × 10 $^{-4}$ mol argon. What is the mass of Ar in a liter of air?

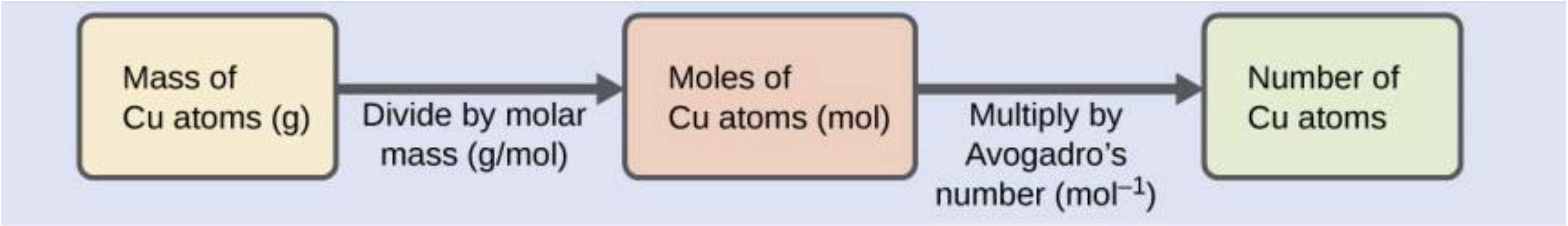


$$9.2\times10^{-4}\ \cancel{\text{mol}}\ \text{Ar}\left(\frac{39.95\ \text{g}}{1\ \cancel{\text{mol}}\ \text{Ar}}\right)=0.037\ \text{g Ar}$$

Figure 3.8

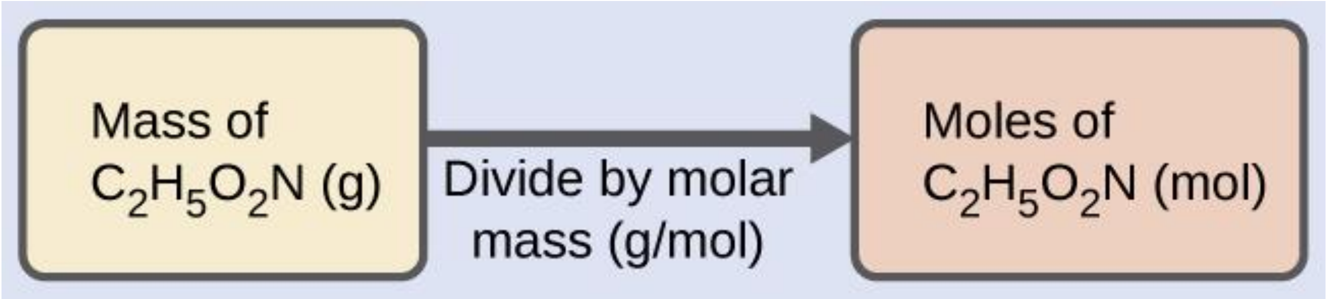
- Copper wire is composed of many, many atoms of Cu. (credit: Emilian Robert Vicol)

Example 3.5



Mass of Cu atoms (g)
Divide by molar mass (g/mol)
Moles of Cu atoms (mol)
Multiply by Avogadro's number (mol^{-1})
Number of Cu atoms

Example 3.6



Mass of $C_2H_5O_2N$ (g)
Divide by molar mass (g/mol)
Moles of $C_2H_5O_2N$ (mol)

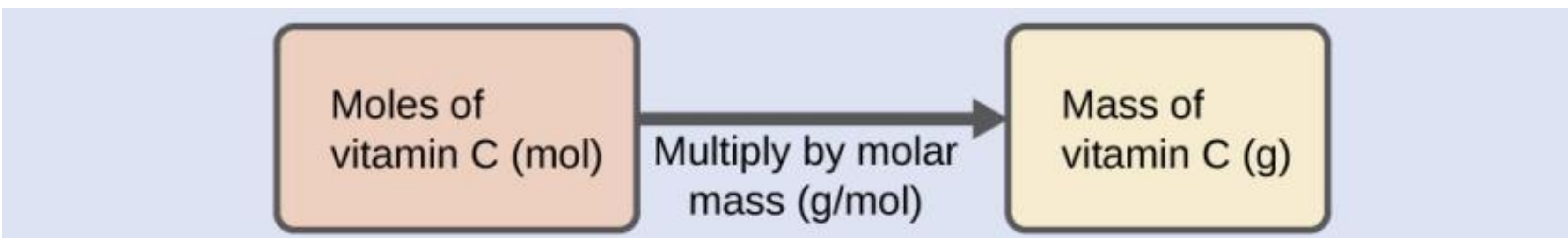
Example 3.6

Element	Quantity (mol element/ mol compound)		Molar mass (g/mol element)		Subtotal (g/mol compound)
C	2	×	12.01	=	24.02
H	5	×	1.008	=	5.040
O	2	×	16.00	=	32.00
N	1	×	14.007	=	14.007
			Molecular mass (g/mol compound)		75.07

$$28.35\ \cancel{\text{g}}\ \text{glycine}\left(\frac{1\ \text{mol glycine}}{75.07\ \cancel{\text{g}}}\right) = 0.378\ \text{mol glycine}$$

Example 3.7

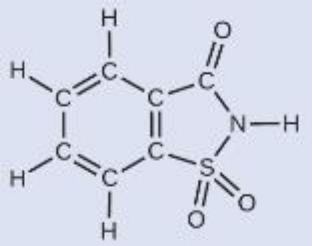
Vitamin C is a covalent compound with the molecular formula $C_6H_8O_6$. The recommended daily dietary allowance of vitamin C for children aged 4–8 years is 1.42×10^{-4} mol. What is the mass of this allowance in grams? The molar mass for this compound is computed to be 176.124 g/mol.



$$1.42 \times 10^{-4}\ \cancel{\text{mol}}\ \text{vitamin C}\left(\frac{176.124\ \text{g}}{1\ \cancel{\text{mol}}\ \text{vitamin C}}\right) = 0.0250\ \text{g vitamin C}$$

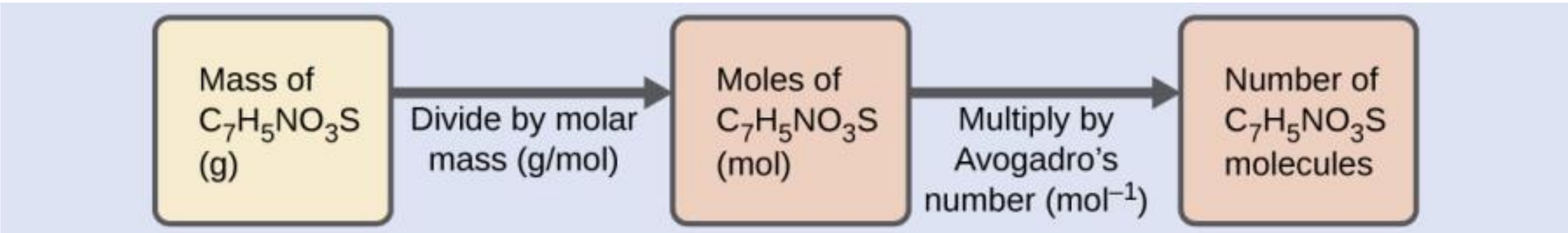
Example 3.8

A packet of an artificial sweetener contains 40.0 mg of saccharin ($C_7H_5NO_3S$), which has the structural formula:



Given that saccharin has a molar mass of 183.18 g/mol, how many saccharin molecules are in a 40.0-mg (0.0400-g) sample of saccharin? How many carbon atoms are in the same sample?

Example 3.8



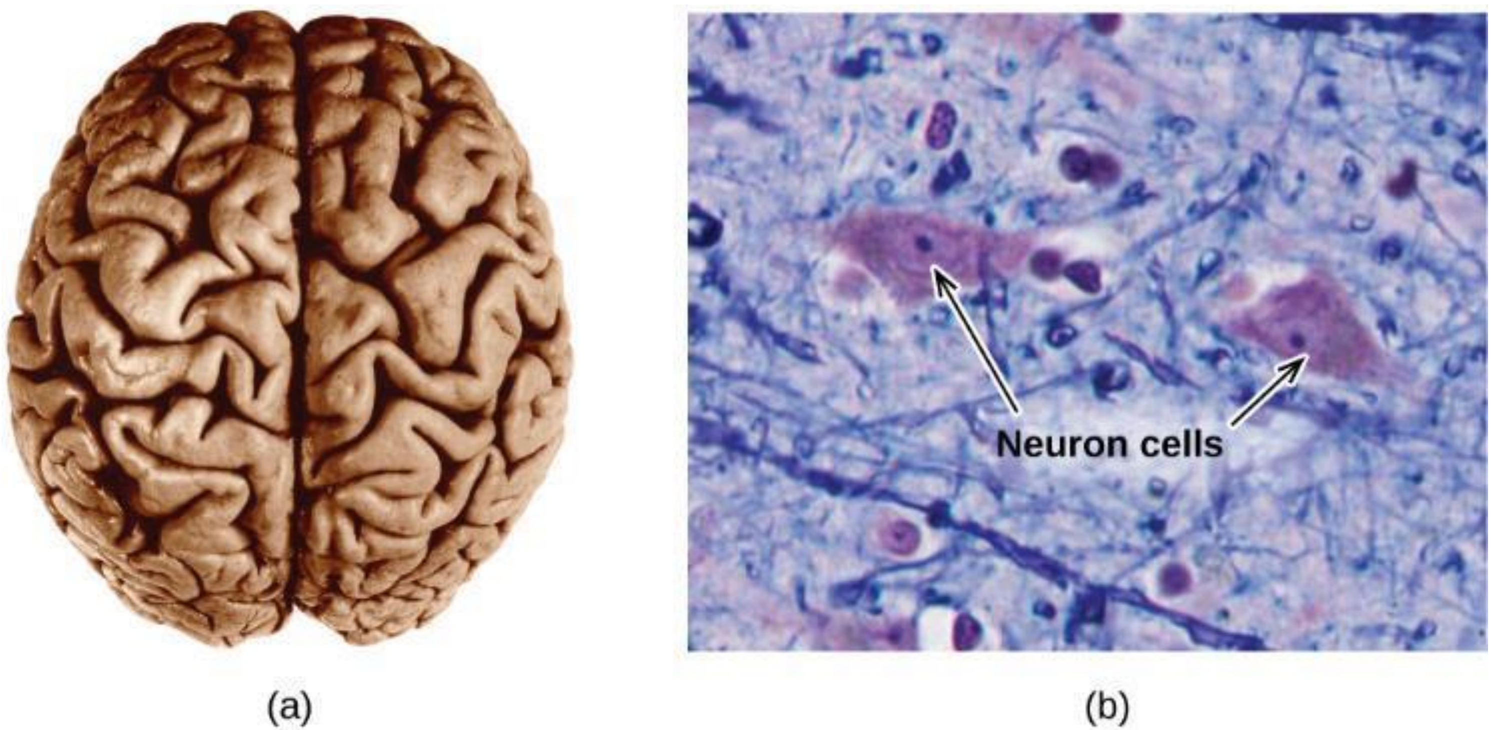
$$0.0400\ \cancel{g}\ C_7H_5NO_3S\left(\frac{1\ \cancel{mol}\ C_7H_5NO_3S}{183.18\ \cancel{g}}\right)\left(\frac{6.022\times10^{23}\ \text{molecules}}{1\ \cancel{mol}\ C_7H_5NO_3S}\right)$$

$$= 1.31\times10^{20}\ C_7H_5NO_3S\ \text{molecules}$$

$$1.31\times10^{20}\ C_7H_5NO_3S\ \text{molecules}\left(\frac{7\ \text{C atoms}}{1\ C_7H_5NO_3S\ \cancel{\text{molecule}}}\right)$$

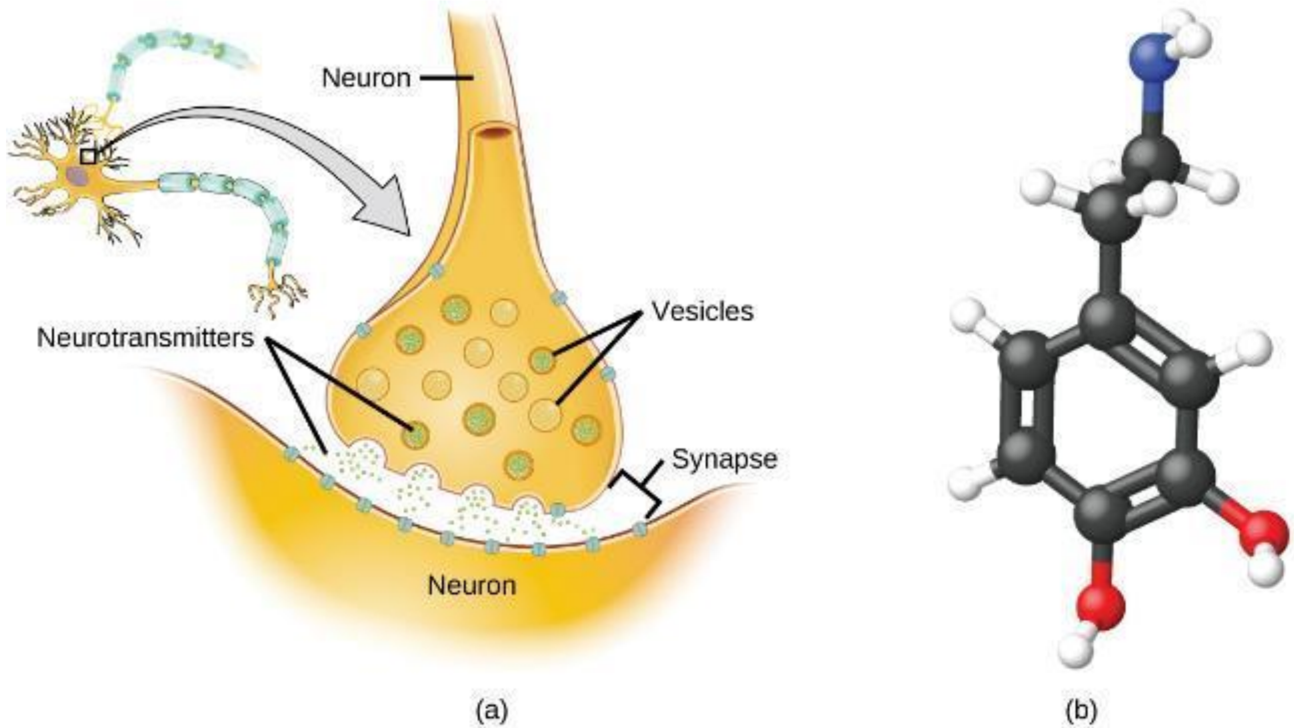
$$= 9.17\times10^{20}\ \text{C atoms}$$

Figure 3.9



(a) A typical human brain weighs about 1.5 kg and occupies a volume of roughly 1.1 L. (b) Information is transmitted in brain tissue and throughout the central nervous system by specialized cells called neurons (micrograph shows cells at 1600× magnification).

Figure 3.10



(a) Chemical signals are transmitted from neurons to other cells by the release of neurotransmitter molecules into the small gaps (synapses) between the cells. (b) Dopamine, $C_8H_{11}NO_2$, is a neurotransmitter involved in a number of neurological processes.

Learning Objectives

- 3.2 Determining Empirical and Molecular Formulas
    - Compute the percent composition of a compound
    - Determine the empirical formula of a compound
    - Determine the molecular formula of a compound

Determining Empirical and Molecular Formulas

- ***Percent composition:*** The percentage by mass of each element in a compound.

- Example: A 10.0 g sample of a compound is determined to contain 2.5 g hydrogen and 7.5 g carbon.

$$\% \text{ H} = \frac{2.5 \text{ g H}}{10.0 \text{ g compound}} \times 100 = 25\%$$

$$\% \text{ C} = \frac{7.5 \text{ g C}}{10.0 \text{ g compound}} \times 100 = 75\%$$

Determining Percent Composition from Molecular or Empirical Formulas

- For compounds of known formula, the percent composition can also be derived from the formula mass and the atomic masses of the constituent elements.

- Example: NH_3 (formula mass = 17.03 amu)

$$\% \text{ N} = \frac{14.01 \text{ amu N}}{17.03 \text{ amu NH}_3} \times 100 = 82.27\%$$

$$\% \text{ H} = \frac{3 \times (1.008 \text{ amu H})}{17.03 \text{ amu NH}_3} \times 100 = 17.76\%$$

Determination of Empirical Formulas

- A compound’s empirical formula can be determined from the masses of its constituent elements.
  1) Convert element masses to moles using molar masses.

  2) Divide each number of moles by the smallest number of moles.

  3) If necessary, multiply by an integer, to give the smallest whole-number ratio of subscripts.

Determination of Empirical Formulas

- Example: A compound is determined to contain 1.71 g C and 0.287 g H.

1)
$$1.17 \text{ g C} \times \frac{1 \text{ mol C}}{12.01 \text{ g C}} = 0.142 \text{ mol C}$$

$$0.287 \text{ g H} \times \frac{1 \text{ mol C}}{1.008 \text{ g H}} = 0.284 \text{ mol H}$$

2)
$$C_{0.142}H_{0.284} = C_{\frac{0.142}{0.142}}H_{\frac{0.284}{0.142}} = CH_2$$

Determination of Empirical Formulas

- **Example:** A compound is determined to contain 5.31 g Cl and 8.40 g O.

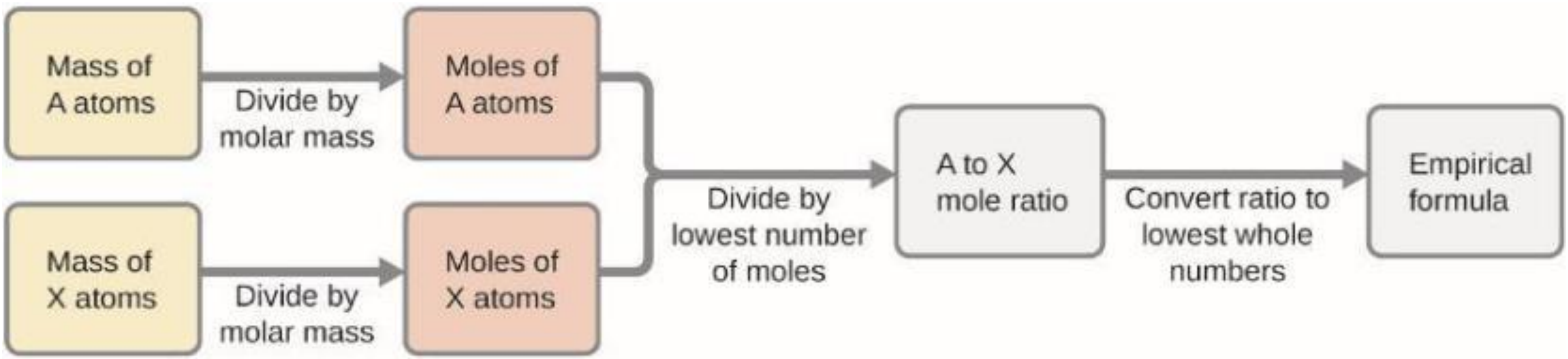
- After steps 1 and 2, the tentative formula is obtained.

$$Cl_{0.150}O_{0.525} = Cl_{\frac{0.150}{0.150}}O_{\frac{0.525}{0.150}} = ClO_{3.5}$$

3) To convert into whole numbers, multiply each of the subscripts by two, giving the empirical formula.

$$Cl_2O_7$$

Figure 3.11



The empirical formula of a compound can be derived from the masses of all elements in the sample.

Determination of Empirical Formulas

- A compound's empirical formula can be determined from its percent composition.
  1) Convert percent composition to masses of elements by assuming a 100 g sample of compound.
  2) Convert element masses to moles using molar masses.
  3) Divide each number of moles by the smallest number of moles.
  4) If necessary, multiply by an integer, to give the smallest whole-number ratio of subscripts.

Figure 3.12

Hematite is an iron oxide that is used in jewelry. (credit: Mauro Cateb)

Figure 3.13

An oxide of carbon is removed from these fermentation tanks through the large copper pipes at the top. (credit: “Dual Freq”/Wikimedia Commons)

Example 3.12

- The bacterial fermentation of grain to produce ethanol forms a gas with a percent composition of 27.29% C and 72.71% O. What is the empirical formula for this gas?

$$27.29\ \text{g C} \times \frac{1\ \text{mol C}}{12.01\ \text{g C}} = 2.272\ \text{mol C}$$

$$72.71\ \text{g O} \times \frac{1\ \text{mol O}}{16.00\ \text{g O}} = 4.544\ \text{mol O}$$

$$\text{C}_{2.272}\text{O}_{4.544} = \text{C}_{\frac{2.272}{2.272}}\text{H}_{\frac{4.544}{2.272}} = \text{CO}_2$$

Derivation of Molecular Formulas

- A compound's molecular formula can be determined from its empirical formula and its molecular or molar mass.

$$\frac{\text{molecular or molar mass}}{\text{empirical formula mass}} = n \text{ formula units/molecule}$$

- The molecular formula is then obtained by multiplying each subscript in the empirical formula by n.

- Example: A compound has an empirical formula of CH_2O (empirical formula mass = 30 amu) and a molecular mass of 180 amu.

$$\frac{180\text{ amu/molecule}}{30\text{ amu/formula unit}} = 6\text{ formula units/molecule}$$

$$\left(\mathrm{CH_2O}\right)_6 = \mathrm{C_6H_{12}O_6}$$

- The molecular formula of this compound is $C_6H_{12}O_6$.

Learning Objectives

- 3.3 Molarity
  - Describe the fundamental properties of solutions
  - Calculate solution concentrations using molarity
  - Perform dilution calculations using the dilution equation

Molarity

- In preceding sections, we focused on the composition of pure substances.
- However, mixtures—samples of matter containing two or more substances physically combined—are more commonly encountered in nature than are pure substances.
- Similar to a pure substance, the relative composition of a mixture plays an important role in determining its properties.

Figure 3.14

Sugar is one of many components in the complex mixture known as coffee. The amount of sugar in a given amount of coffee is an important determinant of the beverage's sweetness. (credit: Jane Whitney)

Solutions

- Solutions occur frequently in nature.
- Solutions are another term used for a **homogeneous mixture—uniform composition and properties throughout its entire volume**.
- The relative amount of a given solution component is known as its **concentration**.

- A solution consists of two components:

1) **Solvent**: component with a concentration that is significantly greater than that of all other components.

2) **Solute**: component that is typically present at a much lower concentration than the solvent.

3) A solution in which water is the solvent is called an **aqueous solution**.

Molarity

- **Molarity (*M*)**: the number of moles of solute in exactly 1 liter (1 L) of the solution:

$$M = \frac{\text{mol solute}}{\text{L solution}}$$

- Molarity (*M*) is a useful concentration unit for many applications in chemistry.

Example 3.14

- A 355 mL soft drink sample contains 0.133 mol of sucrose (table sugar). What is the molar concentration of sucrose in the beverage?

$$M = \frac{\text{mol solute}}{\text{L solution}} = \frac{0.133\ \text{mol}}{355\ \text{mL} \times \dfrac{1\ \text{L}}{1000\ \text{mL}}} = 0.375\ M$$

Figure 3.15

Distilled white vinegar is a solution of acetic acid in water.

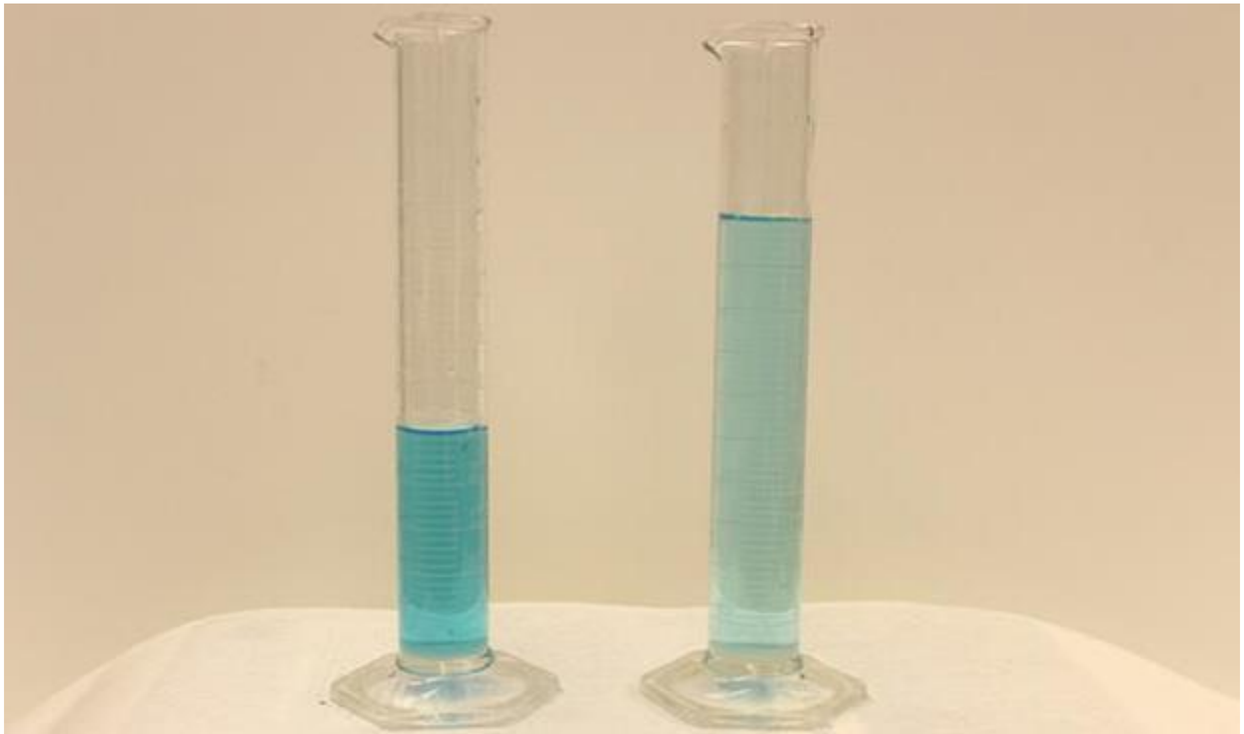
Dilution of Solutions

- **Dilution** is the process whereby the concentration of a solution is lessened by the addition of solvent.

- Dilution is a common means of preparing solutions of a desired concentration.

- By adding solvent to a measured portion of a more concentrated **stock solution**, we can achieve a particular concentration.

Figure 3.16



Both solutions contain the same mass of copper nitrate. The solution on the right is more dilute because the copper nitrate is dissolved in more solvent. (credit: Mark Ott)

Dilution of Solutions

- The molar amount of solute (n) in a solution is equal to the product of the solution's molarity and its volume in liters:

$$n = ML$$

- Expressions like these may be written for a solution **before (1) and after (2)** it is diluted:

$$n_1 = M_1 L_1$$

$$n_2 = M_2 L_2$$

Dilution of Solutions

- Since the dilution process does not change the amount of solute in the solution, $\boldsymbol{n_1 = n_2}$.

- Thus, these two equations may be set equal to one another to derive the **dilution equation**:

$$M_1L_1 = M_2L_2$$

- Other units of concentration (C) and volume (V) may be used.

$$C_1V_1 = C_2V_2$$

Example 3.19

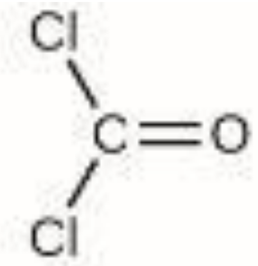
- If 0.850 L of a 5.00-*M* solution of copper nitrate, $Cu(NO_3)_2$, is diluted to a volume of 1.80 L by the addition of water, what is the molarity of the diluted solution?

$$C_1V_1 = C_2V_2$$

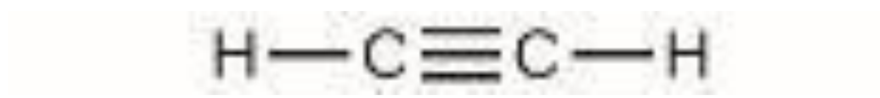
$$C_2 = \frac{C_1V_1}{V_2}$$

$$C_2 = \frac{5.00M \times 0.850\mathrm{L}}{1.80\mathrm{L}} = 2.36M$$

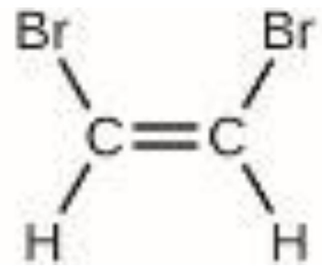
Exercise 4a



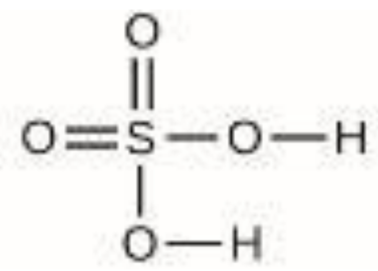
Exercise 4b



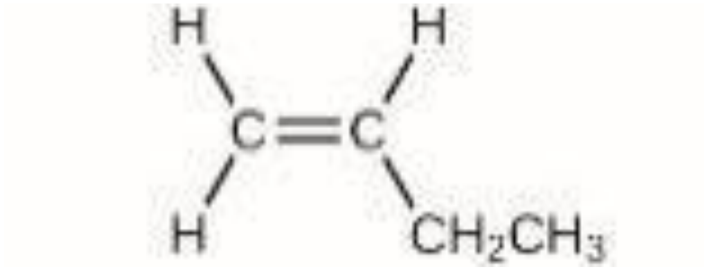
Exercise 4c



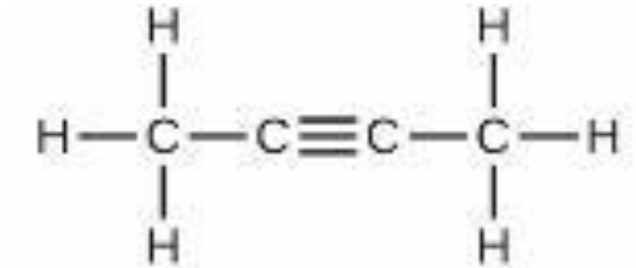
Exercise 4d



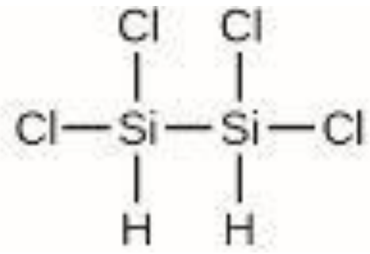
Exercise 5a



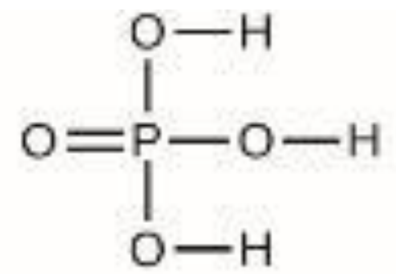
Exercise 5b



Exercise 5c

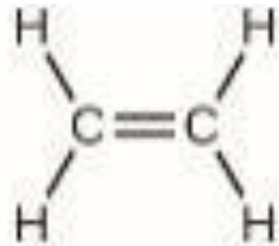


Exercise 5d



Exercise 6a

Exercise 6b



Exercise 6c