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BROOKS/COLE CENGAGE Learning

Chapter 6

Chemical Composition

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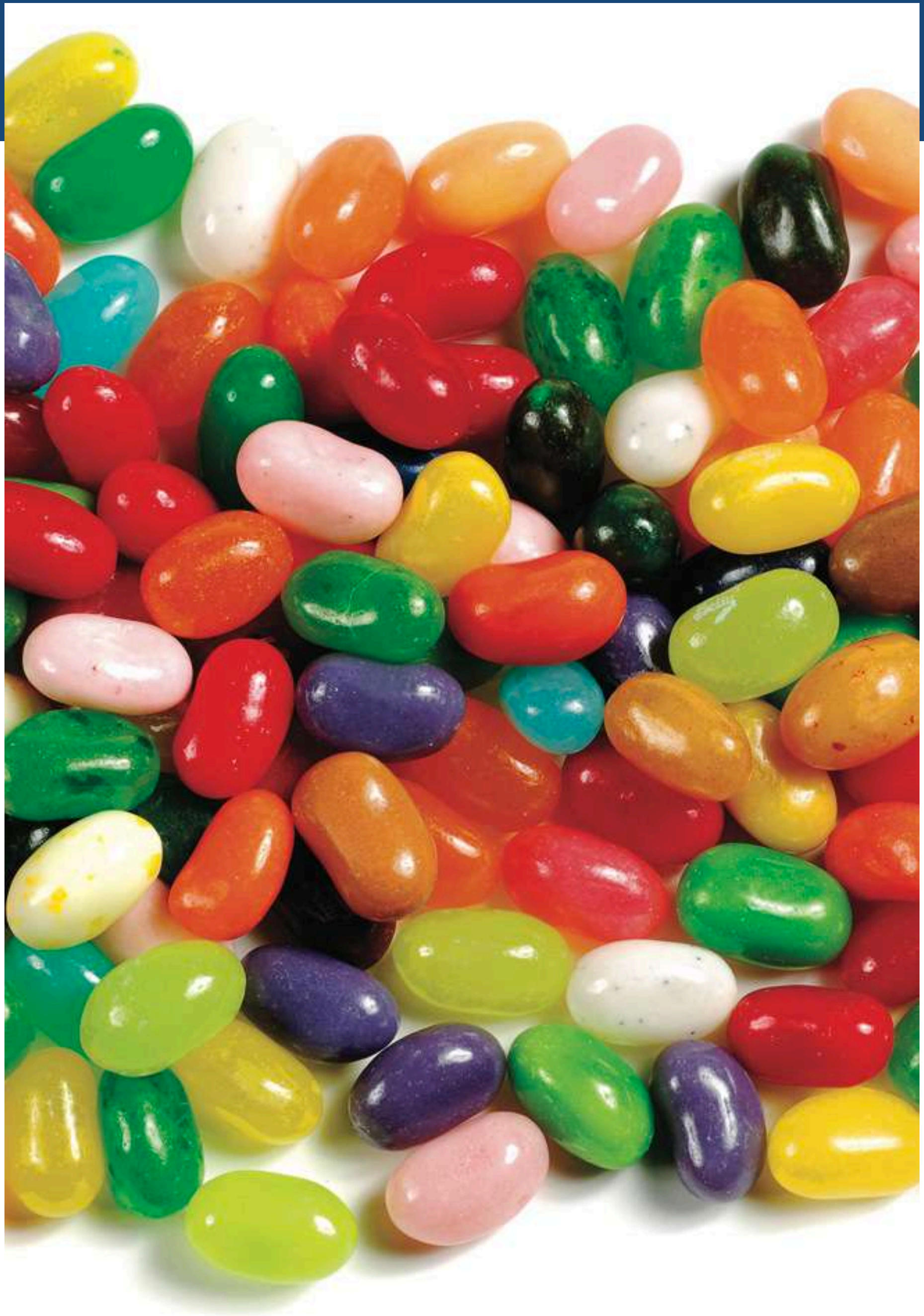
Objectives

1. To understand the concept of average mass
2. To learn how counting can be done by weighing
3. To understand atomic mass and learn how it is determined
4. To understand the mole concept and Avogadro's number
5. To learn to convert among moles, mass, and number of atoms

A. Counting by Weighing

- Objects do not need to have identical masses to be counted by weighing.
  - All we need to know is the average mass of the objects.
- To count the atoms in a sample of a given element by weighing, we must know the mass of the sample and the average mass for that element.

Atoms and Moles



A. Counting by Weighing

Averaging the Mass of Similar Objects

Example: What is the mass of 1000 jelly beans?
Not all jelly beans have the same mass
Suppose we weigh 10 jelly beans and find:

Bean	1	2	3	4	5	6	7	8	9	10
Mass (g)	5.1	5.2	5.0	4.8	4.9	5.0	5.0	5.1	4.9	5.0

Now we can find he average mass of a bean

$$\text{Average mass} = \frac{\text{total mass of beans}}{\text{number of beans}}$$

$$= \frac{5.1\text{ g} + 5.2\text{ g} + 5.0\text{ g} + 4.8\text{ g} + 4.9\text{ g} + 5.0\text{ g} + 5.0\text{ g} + 5.1\text{ g} + 4.9\text{ g} + 5.0\text{ g}}{10}$$

$$= \frac{50.0}{10}\text{ g} = 5.0\text{ g}$$

Finally we can multiply to find the mass of 1000 beans!

A. Counting by Weighing

Averaging the Mass of Different Objects

- Two samples containing different types of components (A and B), both contain the same number of components if the ratio of the sample masses is the same as the ratio of the masses of the individual components.

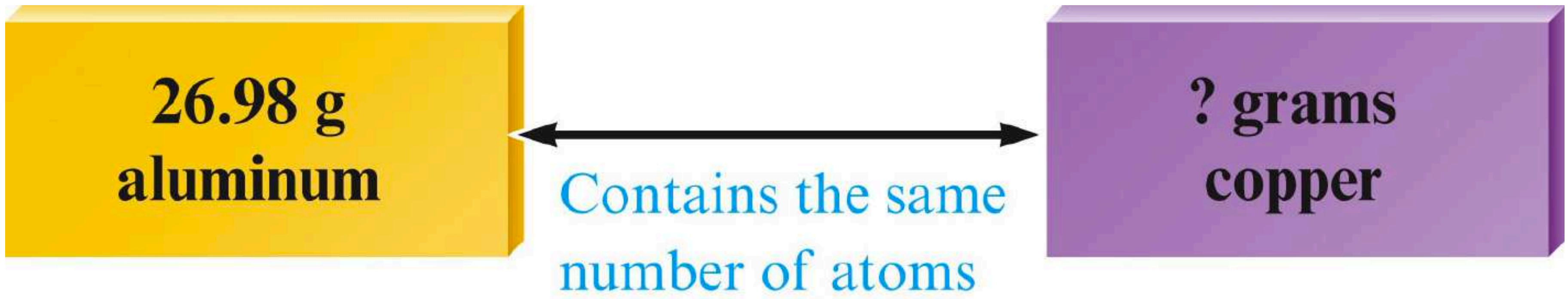
Atoms and Moles

B. Atomic Masses: Counting Atoms by Weighing

- Atoms have very tiny masses so scientists made a unit to avoid using very small numbers.

  1 atomic mass unit (amu) = 1.66 $\times$10–24 g

The average atomic mass for an element is the weighted average of the masses of all the isotopes of an element.


26.98 g
aluminum
Contains the same
number of atoms
? grams
copper

C. The Mole

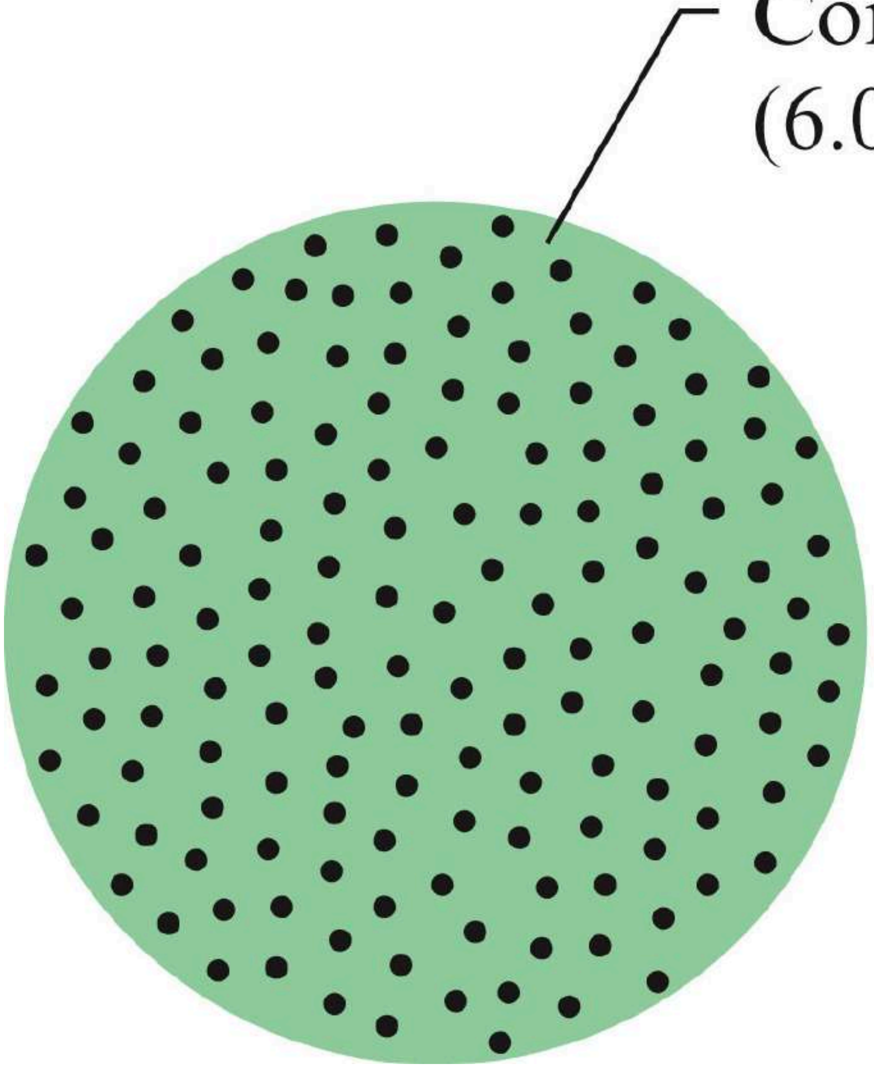
- One mole of anything contains 6.022×10^{23} units of that substance.
  - Avogadro's number is 6.022×10^{23}
- A sample of an element with a mass equal to that element's average atomic mass (expressed in g) contains one mole of atoms.

Average atomic mass of this element	107.9 amu	207.2 amu
Mass of this sample	107.9 g	207.2 g
Number of atoms in this sample	6.022×10^{23} atoms	6.022×10^{23} atoms

C. The Mole

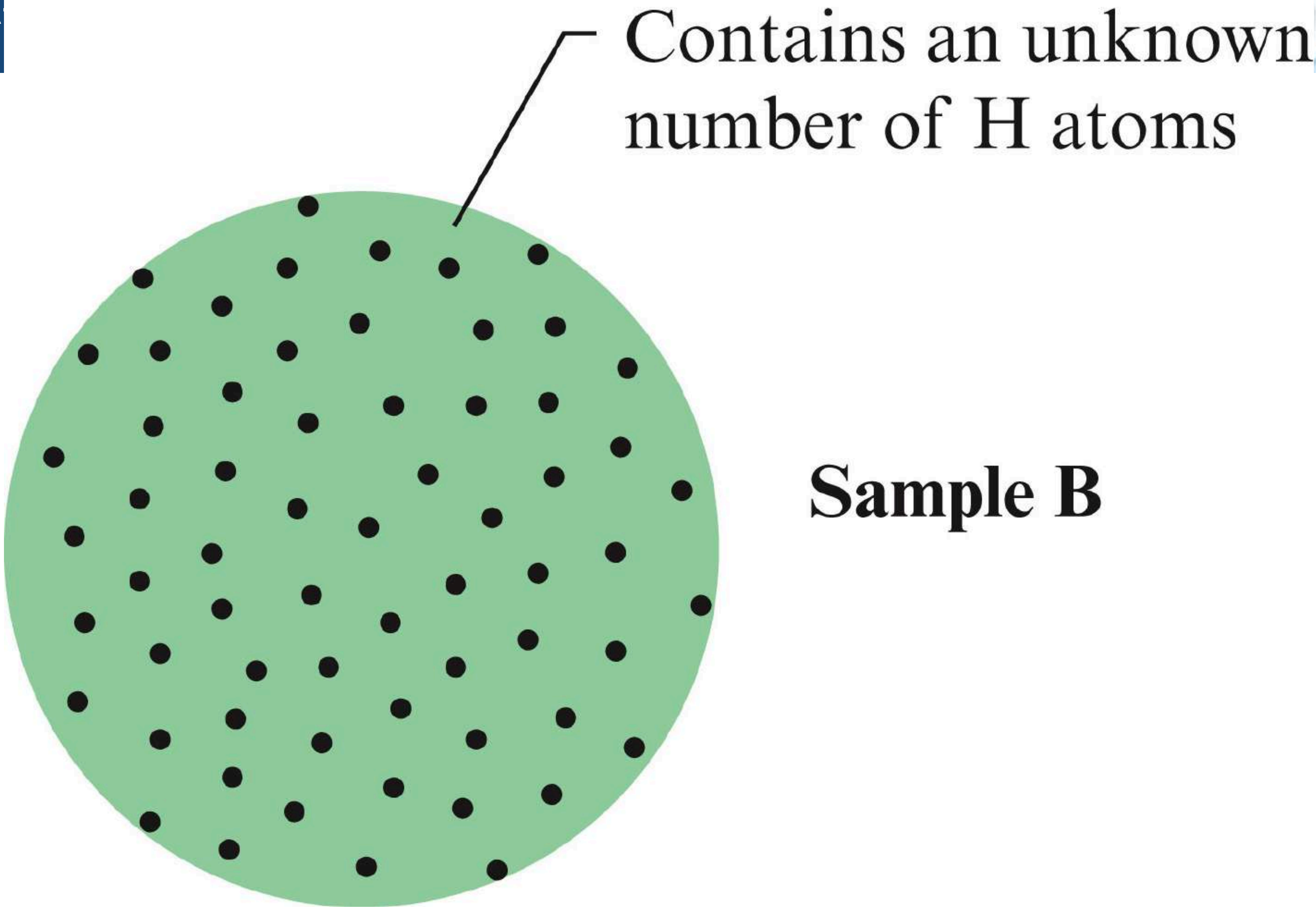
Table 6.1 Comparison of 1-Mol Samples of Various Elements

Element	Number of Atoms Present	Mass of Sample (g)
Aluminum	6.022×10^{23}	26.98
Gold	6.022×10^{23}	196.97
Iron	6.022×10^{23}	55.85
Sulfur	6.022×10^{23}	32.07
Boron	6.022×10^{23}	10.81
Xenon	6.022×10^{23}	131.3



Sample A

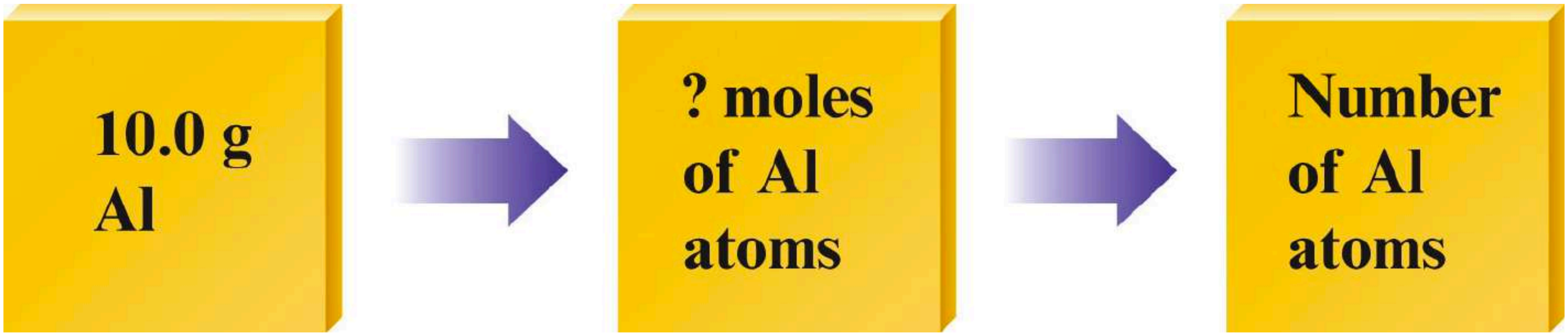
Mass = 1.008 g


Contains an unknown number of H atoms
Sample B

Atoms and Moles

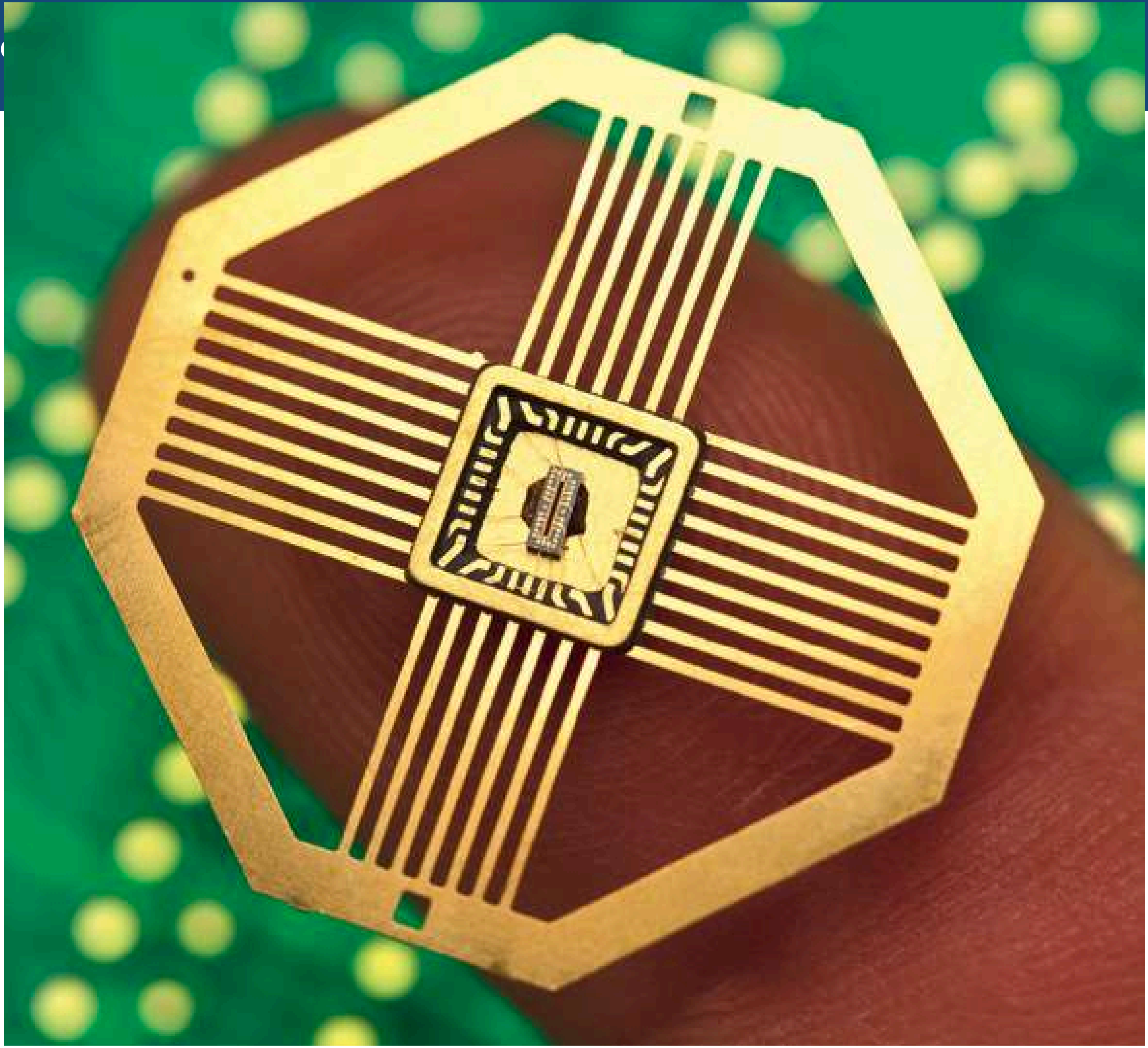
Sample A
Mass = 1.008 g

Sample B
Mass = 0.500 g

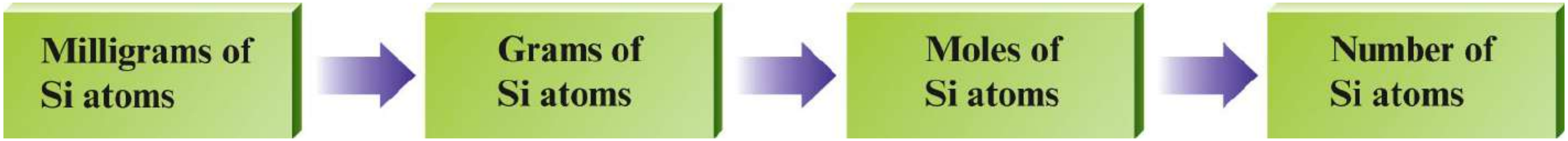

10.0 g Al
? moles of Al atoms
Number of Al atoms

$$10.0 \text{ g Al} \times \frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \rightarrow 0.371 \text{ mol Al}$$

$$0.371 \text{ mol Al} \times \frac{6.022 \times 10^{23} \text{ Al atoms}}{\text{mol Al}} \rightarrow 2.23 \times 10^{23} \text{ Al atoms}$$



5.68 mg Si → Number of Si atoms


Milligrams of
Si atoms
Grams of
Si atoms
Moles of
Si atoms
Number of
Si atoms

Atoms and Moles

$$5.68 \text{ mg Si} \times \frac{1 \text{ g}}{1000 \text{ mg}} \rightarrow 5.68 \times 10^{-3} \text{ g Si}$$

$$5.68 \times 10^{-3} \text{ g Si} \times \frac{1 \text{ mol}}{28.09 \text{ g}} \rightarrow 2.02 \times 10^{-4} \text{ mol Si}$$

$$2.02 \times 10^{-4} \text{ mol Si} \times \frac{6.022 \times 10^{23} \text{ Si atoms}}{\text{mol}} \rightarrow 1.22 \times 10^{20} \text{ Si atoms}$$

Concept Check

Calculate the number of iron atoms in a 4.48 mole sample of iron.

2.70×10^{24} Fe atoms

Concept Check

Which of the following is closest to the average mass of one atom of copper?

a) 63.55 g
b) 52.00 g
c) 58.93 g
d) 65.38 g
e) 1.055×10^{-22} g

Concept Check

Calculate the number of copper atoms in a 63.55 g sample of copper.

6.022×10^{23} Cu atoms

Concept Check

Which of the following 100.0 g samples contains the greatest number of atoms?

a) Magnesium
b) Zinc
c) Silver

Exercise

Rank the following according to number of atoms (greatest to least):

a) 107.9 g of silver

b) 70.0 g of zinc

c) 21.0 g of magnesium

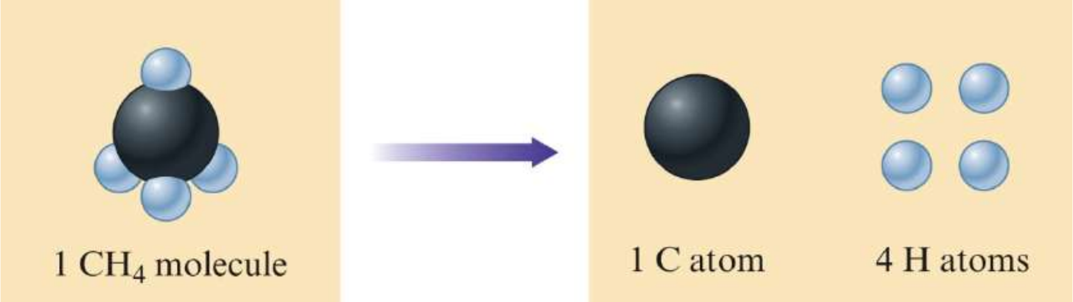
b) a) c)

Objectives

1. To understand the definition of molar mass
2. To learn to convert between moles and mass
3. To learn to calculate the mass percent of an element in a compound

A. Molar Mass

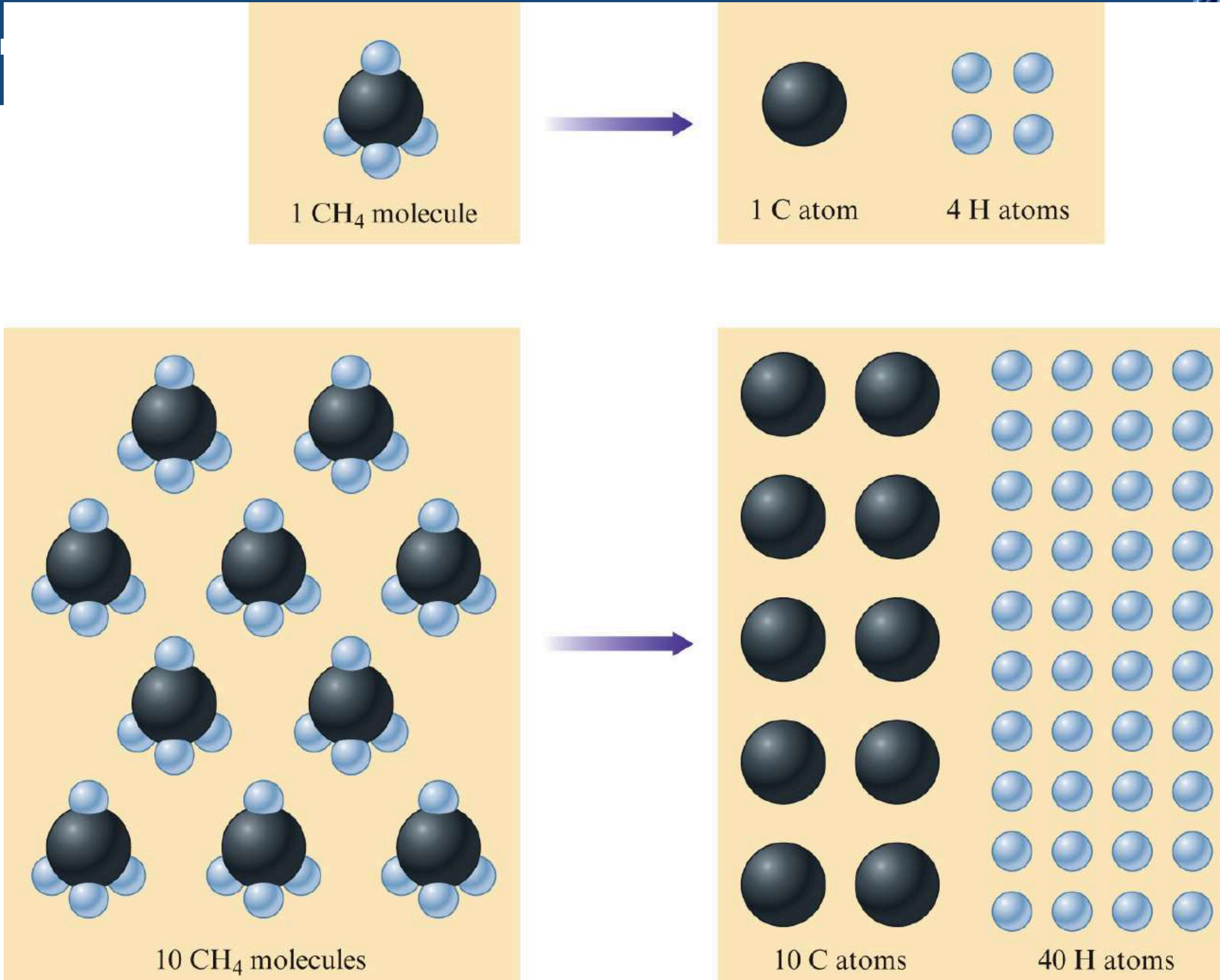
- A compound is a collection of atoms bound together.



The molar mass of a compound is obtained by summing the masses of the component atoms.

$$\begin{aligned} \text{Mass of 1 mol of C} &= 1 \times 12.01\ \text{g} = 12.01\ \text{g} \\ \text{Mass of 4 mol of H} &= 4 \times 1.008\ \text{g} = \underline{4.032\ \text{g}} \\ \text{Mass of 1 mol of } CH_4 &= 16.04\ \text{g} \end{aligned}$$

Molar


1 CH_4 molecule
1 C atom
4 H atoms
10 CH_4 molecules
10 C atoms
40 H atoms

Molar Mass and Percent Composition

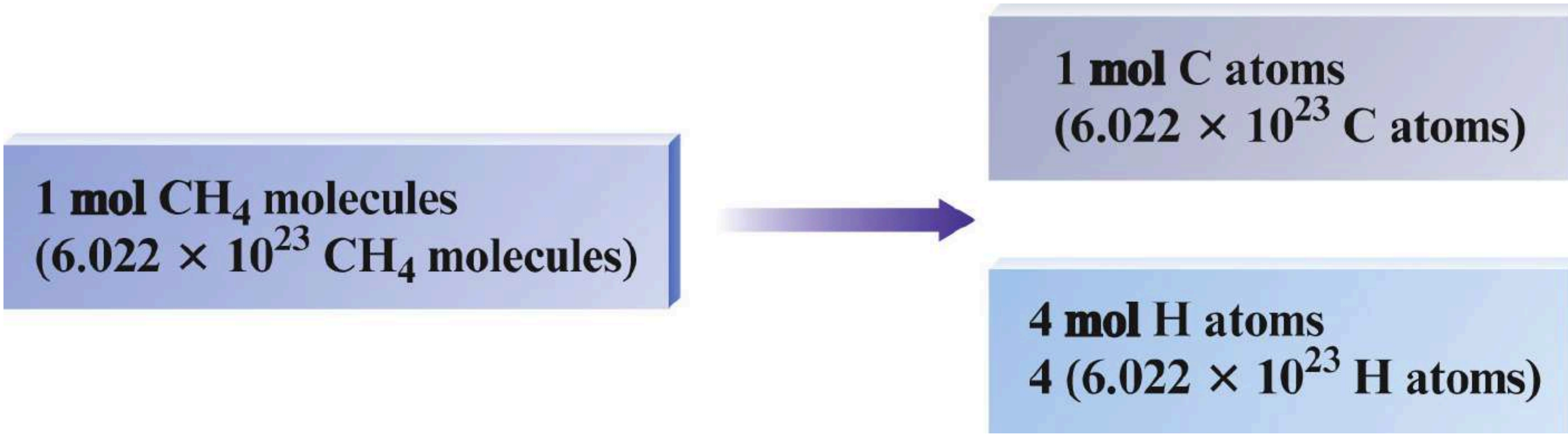
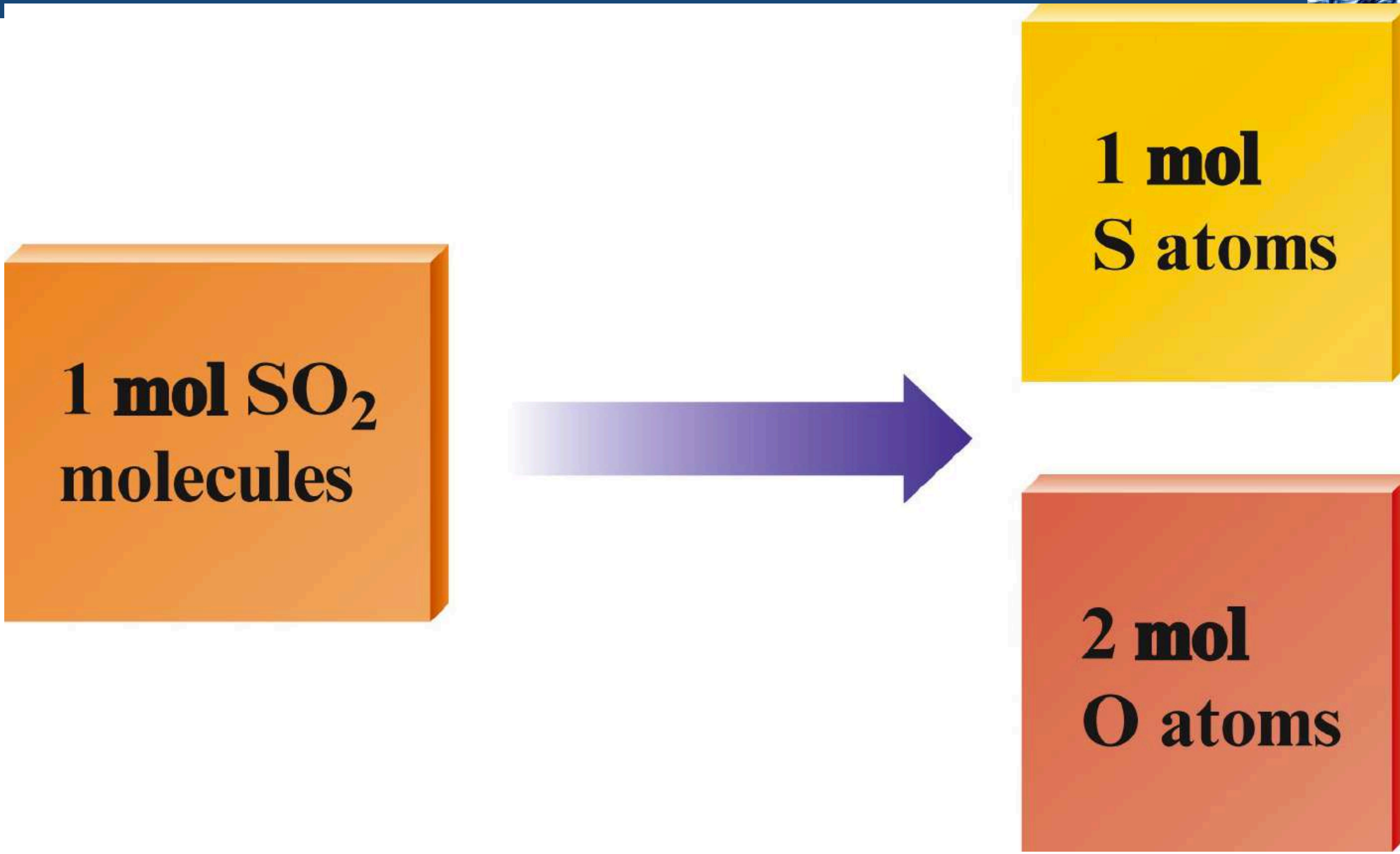
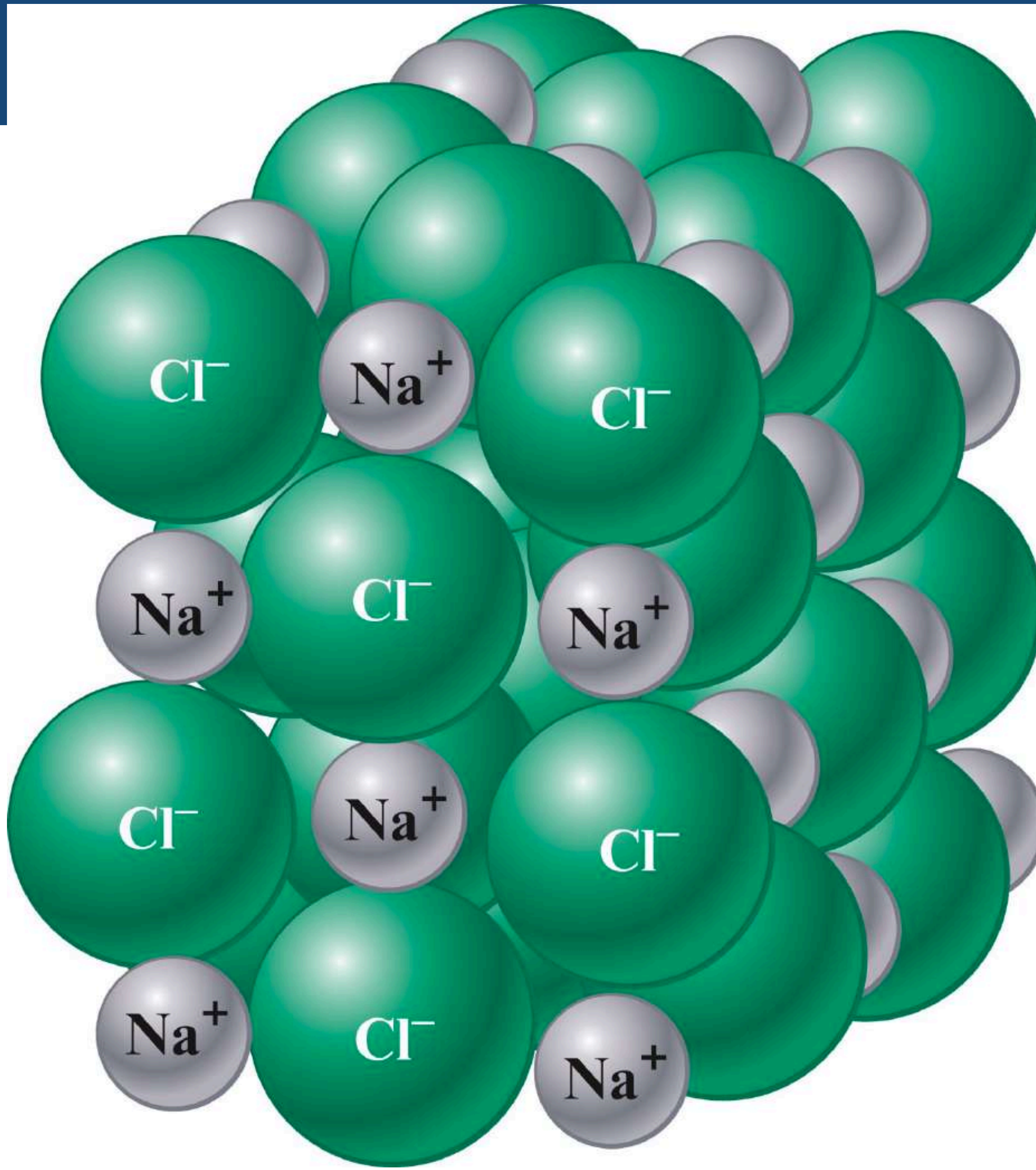


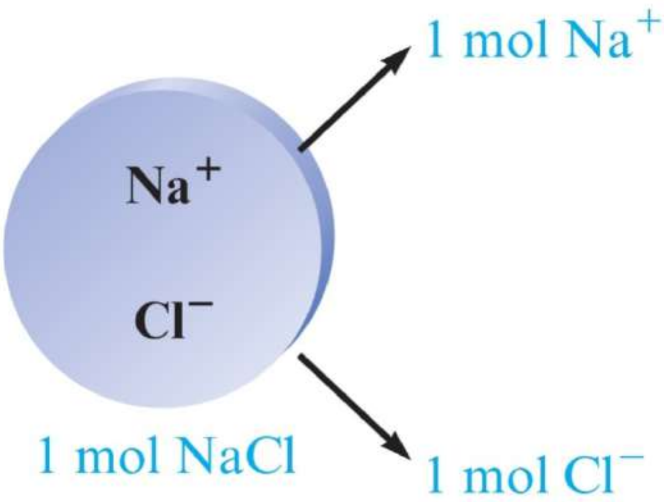
Figure 6-1 p195


1 mol SO_2 molecules
1 mol S atoms
2 mol O atoms


Cl⁻
Na⁺
Cl⁻
Na⁺
Cl⁻
Na⁺
Cl⁻
Na⁺
Cl⁻
Na⁺
Cl⁻
Na⁺

A. Molar Mass

- For compounds containing ions, the molar mass is obtained by summing the masses of the component ions.



$$\begin{aligned}\text{Mass of 1 mol } Na^+ &= 22.99\text{ g}\\ \text{Mass of 1 mol } Cl^- &= \underline{35.45\text{ g}}\\ \text{Mass of 1 mol } NaCl &= 58.44\text{ g} = \text{molar mass}\end{aligned}$$

- Mass in grams of one mole of the substance:

Molar Mass of N = 14.01 g/mol

Molar Mass of H_2O = 18.02 g/mol

(2 × 1.008 g) + 16.00 g

Molar Mass of $Ba(NO_3)_2$ = 261.35 g/mol

137.33 g + (2 × 14.01 g) + (6 × 16.00 g)

A. Molar Mass

Calculations Using Molar Mass

- Moles of a compound = $\dfrac{\text{mass of the sample (g)}}{\text{molar mass of the compound }(\frac{\text{g}}{\text{mol}})}$

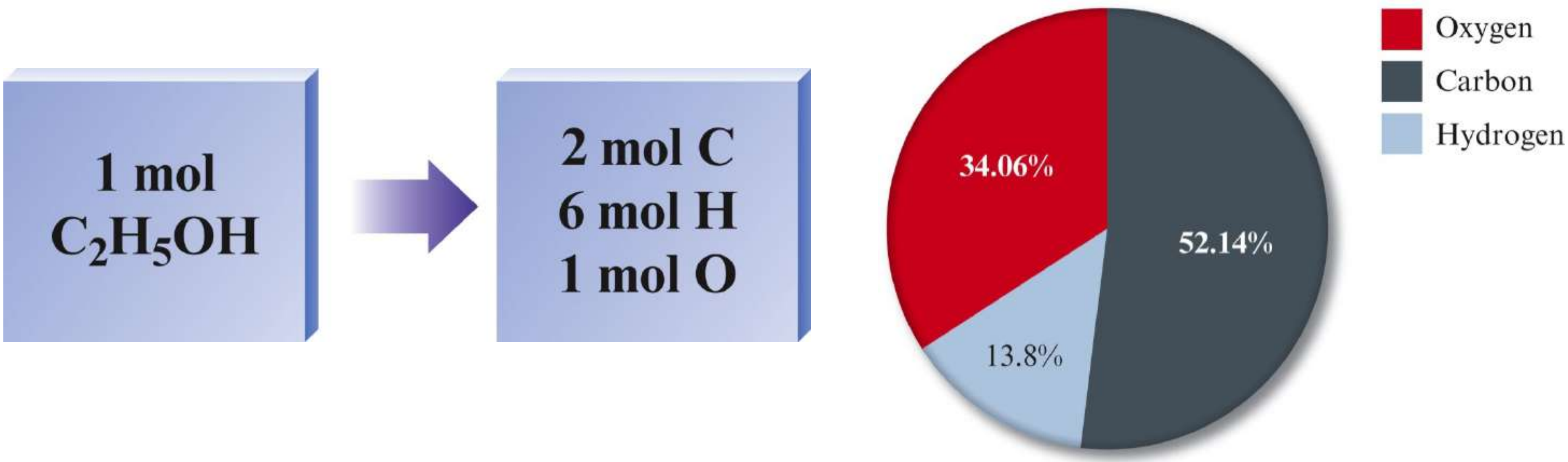
Mass of a sample (g) = (moles of sample)(molar mass of compound)

$$4.86 \text{ mol } CaCO_3 \times \frac{100.09 \text{ g}}{\text{mol}} \rightarrow 486 \text{ g } CaCO_3$$

B. Percent Composition of Compounds

- Percent composition consists of the mass percent of each element in a compound:

Mass percent = $\dfrac{\text{mass of a given element in 1 mol of compound}}{\text{mass of 1 mol of compound}} \times 100\%$



- For iron in iron(III) oxide,(Fe_2O_3):

$$\text{mass \% Fe} = \frac{2(\ \ 55.85\text{ g})}{2(\ \ 55.85\text{ g}) + 3(\ \ 16.00\text{ g})} = \frac{111.70\text{ g}}{159.70\text{ g}} \times 100\% = 69.94\%$$

Exercise

Consider separate 100.0 gram samples of each of the following:

H_2O, N_2O, $C_3H_6O_2$, CO_2

- Rank them from highest to lowest percent oxygen by mass.

H_2O, CO_2, $C_3H_6O_2$, N_2O

Objectives

1. To understand the meaning of empirical formula
2. To learn to calculate empirical formulas
3. To learn to calculate the molecular formula of a compound

A. Empirical Formulas

- The empirical formula of a compound is the simplest whole number ratio of the atoms present in the compound.
- The empirical formula can be found from the percent composition of the compound.

B. Calculation of Empirical Formulas

Tools for Determining the Empirical Formula of a Compound

Step 1 Obtain the mass of each element present (in grams).

Step 2 Determine the number of moles of each type of atom present.

Step 3 Divide the number of moles of each element by the smallest number of moles to convert the smallest number to 1. If all of the numbers so obtained are whole numbers, they are the subscripts in the empirical formula. If one or more of these numbers are not whole numbers, go on to step 4.

Step 4 Multiply the numbers from step 3 by the smallest whole number that will convert all of them to whole numbers. This set of whole numbers represents the subscripts in the empirical formula.

C. Calculation of Molecular Formulas

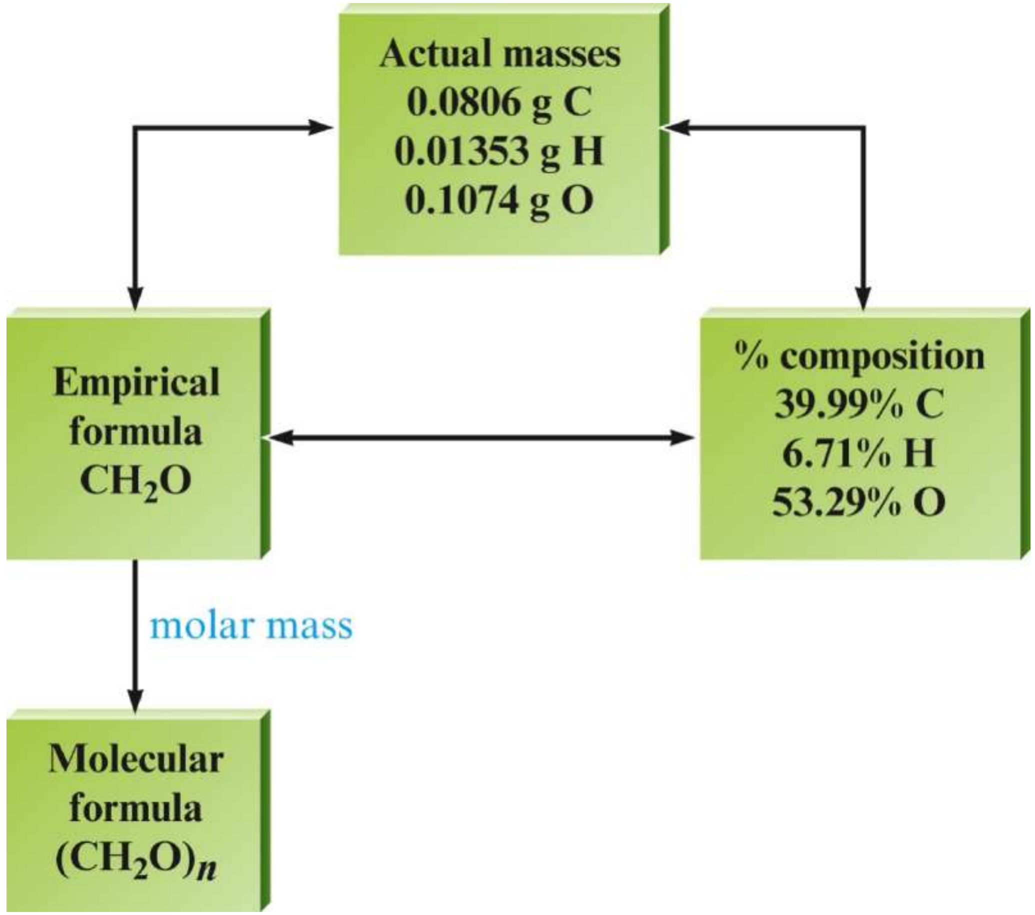
- The molecular formula is the exact formula of the molecules present in a substance.
- The molecular formula is always an integer multiple of the empirical formula.

Molecular formula = (empirical formula)n

where n is a whole number

Formula Summary

For the sugar glucose



Exercise

The composition of adipic acid is 49.3% C, 6.9% H, and 43.8% O (by mass). The molar mass of the compound is about 146 g/mol.

- What is the empirical formula?

$C_3H_5O_2$

- What is the molecular formula?

$C_6H_{10}O_4$