Chapter 3. Stoichiometry: Mole-Mass Relationships in Chemical Reactions

Concept 1. The meaning and usefulness of the mole

- The mole (or mol) represents a certain number of objects.
- SI def.: the amount of a substance that contains the same number of entities as there are atoms in 12 g of carbon-12.
- Exactly 12 g of carbon-12 contains **6.022×10^{23}** atoms.
- 1 mole contains **6.022×10^{23}** entities (Avogadro's number)
- One mole of H_2O molecules contains 6.022×10^{23} molecules.

- One mole of NaCl contains 6.022×10^{23} NaCl formula units.
- Use the mole quantity to count formulas by weighing them.
- Mass of a mole of particles = mass of 1 particle **x** 6.022×10^{23}

Mass of 1 H atom: 1.008 amu x 1.661×10^{-24} g/amu = 1.674×10^{-24} g

Mass of 1 mole of H atoms:

1.674×10^{-24} g/H atom x 6.022×10^{23} H atoms = 1.008 g

- The mass of an atom in amu is numerically the same as the mass of one mole of atoms of the element in grams.
- One atom of sulfur has a mass of 32.07 amu; one mole of S atoms has a mass of 32.07 g.

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Concept 2. The relation between molecular (formula) mass and molar mass

- For compounds, the molecular mass (in amu) is numerically the same as the mass of one mole of the compound in grams.

- **Skill 3-1** Calculate the molecular mass of a compound as the sum of the atomic masses of its elements.

- **Molecular mass H_2O = (2 x atomic mass of H) + atomic mass of O**

  = 2(1.008 amu) + 16.00 amu = 18.02 amu

- **So, one mole of water (6.022 x 10^{23} molecules) has a mass of 18.02 g.**

- Molar mass of NaCl = atomic mass of Na (22.99 amu) + the atomic mass of Cl (35.45 amu) **22.99** + **35.45** = **58.44** amu

- One mol of NaCl (6.02 x10^{23} formulas) has a mass of 58.44 g.

- To obtain one mole of copper atoms (6.02×10^{23} atoms), weigh out 63.55 g copper.

Concept 3. The relations among amount of substance (in moles), mass (in grams), and number of chemical entities

- The molar mass (M) of a substance is the mass of one mole of entities (atoms, molecules, or formula units) of the substance.
- Molar mass has units of grams per mole (g/mol).

TABLE 3.1 Summary of Mass Terminology*

TERM	DEFINITION	UNIT
Isotopic mass	Mass of an isotope of an element	amu
Atomic mass (also called atomic weight)	Average of the masses of the naturally occurring isotopes of an element weighted according to their abundance	amu
Molecular mass (also called molecular weight)	Sum of the atomic masses of the atoms (or ions) in a molecule (or formula unit)	amu
Molar mass ($\mathcal{M}$) (also called gram-molecular weight)	Mass of one mole of chemical entities (atoms, ions, molecules, formula units)	g/mol

Skill 3-2 Mass - Mole Conversions

- **Use the molar mass of an element or compound to convert a given number of moles to mass:**

$$\text{Mass (g)} = \text{no. of moles} \times \frac{\text{no. of grams}}{1 \text{ mole}}$$

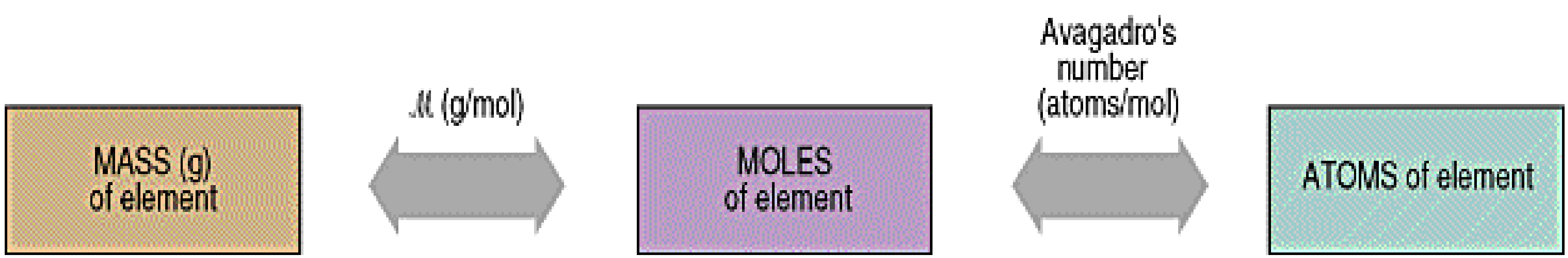
- **We can do the reverse with 1/*M*, and convert any mass in grams to the number of moles:**

$$\text{No. of moles} = \text{mass (g)} \times \frac{1 \text{ mole}}{\text{no. of grams}}$$

- **Use Avogadro's number to convert moles of substance to the number of entities:**

$$\text{No. of entities} = \text{no. of moles} \times \frac{6.022 \times 10^{23} \text{ entities}}{1 \text{ mole}}$$

$$\text{No. of moles} = \text{no. of entities} \times \frac{1 \text{ mole}}{6.022 \times 10^{23} \text{ entities}}$$



Problem: (a) How many grams of silver, Ag, are in 0.0342 mol Ag? (b) How many atoms of Ag are in 0.0342 mol Ag?

Plan: (a) To convert moles Ag to grams Ag, use the molar mass of Ag from the periodic table. (b) To convert moles to number of atoms, use Avogadro's number.

Solution: (a) Converting from moles of Ag to mass:

$$\text{Mass (g) of Ag} = 0.0342\ \cancel{\text{mol Ag}} \times \frac{107.9\ \text{g Ag}}{1\ \cancel{\text{mol Ag}}} = \mathbf{3.69\ g\ Ag}$$

(b) Converting from moles of Ag to number of atoms:

$$\text{No. of Ag atoms} = 0.0342\ \cancel{\text{mol Ag}} \times \frac{6.022 \times 10^{23}\ \text{atoms Ag}}{1\ \cancel{\text{mol Ag}}}$$

$$= 2.06 \times 10^{22}\ \text{atoms Ag}$$

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Problem: Iron is a most important metal in our society. How many iron atoms are present in a piece of iron weighing 95.8 g?

Plan:

$$\text{Mass (g) of Fe} \xrightarrow{\mathcal{M}\text{ of Fe} = 55.85\text{ g/mol}} \text{Moles of Fe} \xrightarrow{6.022 \times 10^{23}\text{ atoms/mol}} \text{Number of Fe atoms}$$

Solution: Converting from mass of Fe to moles:

$$\text{Moles of Fe} = 95.8\ \cancel{\text{g Fe}} \times \frac{1\text{ mol Fe}}{55.85\ \cancel{\text{g Fe}}} = 1.72\text{ mol Fe}$$

Converting from moles of Fe to number of atoms:

$$\text{No. of Fe atoms} = 1.72\ \cancel{\text{mol Fe}} \times \frac{6.022 \times 10^{23}\text{ atoms Fe}}{1\ \cancel{\text{mol Fe}}}$$

$$= 10.4 \times 10^{23}\text{ atoms Fe} = \mathbf{1.04 \times 10^{24}\text{ atoms Fe}}$$

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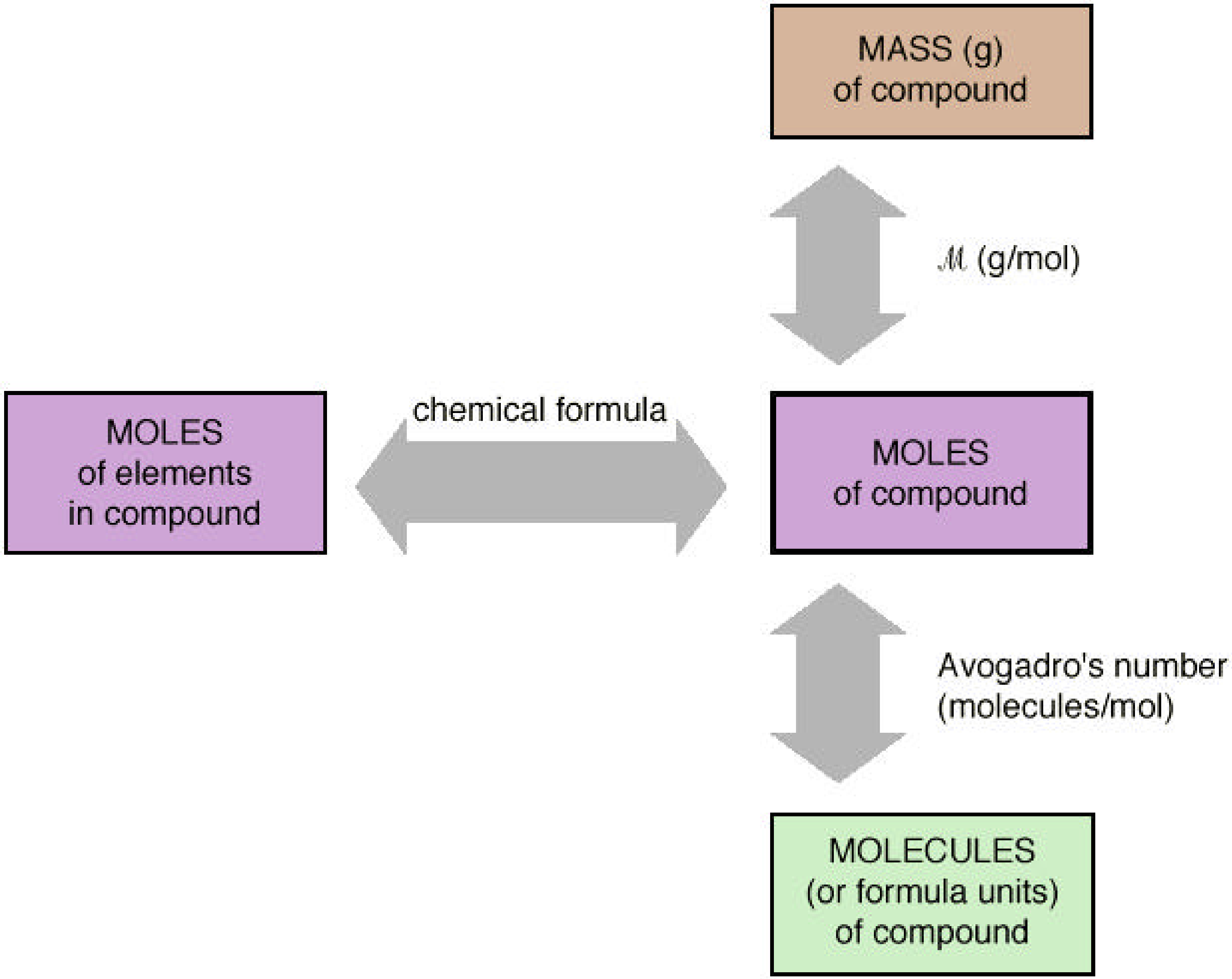
Concept 4. The information in a chemical formula

- The sulfur dioxide formula, SO_2, shows that its molecules contain one S atom and two O atoms; calculate its molar mass.

- One mole SO_2 contains 6.02×10^{23} SO_2 molecules, which consist of 6.02×10^{23} S atoms and $2(6.02 \times 10^{23})$ O atoms.

- Same for ionic compounds, such as potassium sulfide (K_2S):


MASS (g)
of compound
$\mathcal{M}$ (g/mol)
MOLES
of elements
in compound
chemical formula
MOLES
of compound
Avogadro's number
(molecules/mol)
MOLECULES
(or formula units)
of compound

Problem: How many moles and formulas are in 41.6 g ammonium carbonate?

$$\text{Mass (g) of } (NH_4)_2CO_3 \xrightarrow{\mathcal{M}\ (\text{g/mol})} \text{Moles of } (NH_4)_2CO_3 \xrightarrow{6.022 \times 10^{23} \text{ formula units/mol}} \text{Number of } (NH_4)_2CO_3 \text{ formula units}$$

Solution: Calculating molar mass:

$$\mathcal{M} = (2 \times 14.01 \text{ g/mol}) + (8 \times 1.008 \text{ g/mol}) + 12.01 \text{ g/mol} + (3 \times 16.00 \text{ g/mol})$$

$$= 96.09 \text{ g/mol}$$

Converting from mass to moles:

$$\text{Moles of } (NH_4)_2CO_3 = 41.6 \cancel{\text{g } (NH_4)_2CO_3} \times \frac{1 \text{ mol } (NH_4)_2CO_3}{96.09 \cancel{\text{g } (NH_4)_2CO_3}}$$

$$= \mathbf{0.433 \text{ mol } (NH_4)_2CO_3}$$

Converting from moles to formula units:

Formula units of $(NH_4)_2CO_3$

$$= 0.433 \cancel{\text{mol } (NH_4)_2CO_3} \times \frac{6.022 \times 10^{23} \text{ formula units } (NH_4)_2CO_3}{1 \cancel{\text{mol } (NH_4)_2CO_3}}$$

$$= \mathbf{2.61 \times 10^{23} \text{ formula units } (NH_4)_2CO_3}$$

Skill 3-3 Mass Percent and the Chemical Formula

The formula shows the number of moles of each element; use it to calculate the mass percent of each element on a mole basis:

$$\text{Mass \% of element X} = \frac{\text{moles of X} \times \text{molar mass of X (g/mol)}}{\text{mass of one mole of compound (g)}} \times 100\ \%$$

Problem: The formula of the sugar glucose is $\mathbf{C_6H_{12}O_6}$.

(a) What is the mass percent of each element in glucose?

(b) How many grams of carbon are in 16.55 g glucose?

Plan: The formula gives the no. of moles of C, H and O in glucose. Convert from moles of element to grams using the molar mass.

$$\text{Mass (g) of C} = 6\ \cancel{\text{mol C}} \times \frac{12.01\ \text{g C}}{1\ \cancel{\text{mol C}}} = 72.06\ \text{g C}$$

Dividing by the mass of one mole of glucose gives the element's mass fraction; multiplying this fraction by 100% gives the mass percent.

$$\text{Mass fraction of C} = \frac{\text{total mass C}}{\text{mass of 1 mol glucose}} = \frac{72.06\text{ g}}{180.16\text{ g}} = 0.4000\text{ g C/g glucose}$$

$$\text{Mass \% of C} = \text{mass fraction of C} \times 100 = 0.4000 \times 100$$

$$= \mathbf{40.00\ mass\ \%\ C}$$

Combining the steps for each of the other two elements:

$$\text{Mass \% of H} = \frac{\text{mol H} \times \mathcal{M}\text{ of H}}{\text{mass of 1 mol glucose}} \times 100 = \frac{12\ \cancel{\text{mol H}} \times \dfrac{1.008\ \cancel{\text{g}}\text{ H}}{1\ \cancel{\text{mol H}}}}{180.16\ \cancel{\text{g}}} \times 100$$

$$= \mathbf{6.714\ mass\ \%\ H}$$

$$\text{Mass \% of O} = \frac{\text{mol O} \times \mathcal{M}\text{ of O}}{\text{mass of 1 mol glucose}} \times 100 = \frac{6\ \cancel{\text{mol O}} \times \dfrac{16.00\ \cancel{\text{g}}\text{ O}}{1\ \cancel{\text{mol O}}}}{180.16\ \cancel{\text{g}}} \times 100$$

$$= \mathbf{53.29\ mass\ \%\ O}$$

(b) Determining the mass of carbon.

Plan: To find the mass of C in the glucose sample, multiply the mass of the sample by the mass fraction of C from part (a)

Solution:

$$\text{Mass (g) of C} = \text{mass of glucose} \times \text{mass fraction of C}$$

$$= 16.55\ \cancel{\text{g glucose}} \times \frac{0.4000\ \text{g C}}{1\ \cancel{\text{g glucose}}} = \mathbf{6.620\ g\ C}$$

Concept 5. The procedure for finding the empirical and molecular formulas of a compound

- If an unknown pure compound can be **decomposed**, the masses of elements present can be determined.

- Converting the masses to moles of elements leads to the empirical formula, the simplest whole-number ratio of moles of each element in the compound.

- Suppose a sample of unknown compound is found to contain **0.21 mol zinc**, **0.14 mol phosphorus**, and **0.56 mol oxygen**.

- A preliminary formula based on the data is **$Zn_{0.21}P_{0.14}O_{0.56}$**.
- Convert the fractional subscripts to whole numbers. :

1. Divide each subscript by the smallest subscript:

$$Zn_{\frac{0.21}{0.14}}P_{\frac{0.14}{0.14}}O_{\frac{0.56}{0.14}} \rightarrow Zn_{1.5}P_{1.0}O_{4.0}$$

2. If any subscripts are still not integers, multiply through by the smallest integer that will turn all the subscripts into integers

$$Zn_{(1.5\times 2)}P_{(1.0\times 2)}O_{(4.0\times 2)} \rightarrow Zn_{3.0}P_{2.0}O_{8.0}$$

$$\text{or } Zn_3P_2O_8$$

The conventional way to write this formula is **$Zn_3(PO_4)_2$** ; the compound is zinc phosphate, a dental cement.

Skill 3-4: Analysis of a sample of ionic compound gave: 2.82 g Na, 4.35 g Cl, and 7.83 g O. What is the empirical formula and name of the compound?

Plan:

Mass (g) of element $\xrightarrow{\mathcal{M}\ (\text{g/mol})}$ Moles of element $\xrightarrow{\text{use nos. of moles as subscripts}}$ Preliminary formula $\xrightarrow{\text{change to integer subscripts}}$ Empirical formula

Solution: Find moles of elements:

$$\text{Moles of Na} = 2.82\ \cancel{\text{g Na}} \times \frac{1\ \text{mol Na}}{22.99\ \cancel{\text{g Na}}} = 0.123\ \text{mol Na}$$

$$\text{Moles of Cl} = 4.35\ \cancel{\text{g Cl}} \times \frac{1\ \text{mol Cl}}{35.45\ \cancel{\text{g Cl}}} = 0.123\ \text{mol Cl}$$

$$\text{Moles of O} = 7.83\ \cancel{\text{g O}} \times \frac{1\ \text{mol O}}{16.00\ \cancel{\text{g O}}} = 0.489\ \text{mol O}$$

From the moles of each element, construct a preliminary formula and convert to integer subscripts.

Preliminary formula: $Na_{0.123}Cl_{0.123}O_{0.489}$

Convert to integer subscripts:

$$Na_{\frac{0.123}{0.123}}Cl_{\frac{0.123}{0.123}}O_{\frac{0.489}{0.123}} \longrightarrow Na_{1.00}Cl_{1.00}O_{3.98} \approx Na_1Cl_1O_4, \text{ or } NaClO_4$$

- Note that we rounded the subscript of O from 3.98 to 4.

- The empirical formula is $NaClO_4$; the name is sodium perchlorate.

Molecular Formulas

- The actual number of moles of each element in the smallest unit of the compound.
- In water (H_2O), ammonia (NH_3), methane (CH_4), and ionic compounds, the empirical and molecular formulas are identical.
- In some cases the molecular formula is a whole-number multiple of the empirical formula.

- Using the empirical formula to obtain the molecular formula.

Hydrogen peroxide has the empirical formula **HO** (17.01g/mol)

Dividing its molar mass (34.02 g/mol) by the empirical formula mass gives the whole-number multiple:

$$\text{Whole-number multiple} = \frac{\text{molar mass (g/mol)}}{\text{empirical formula mass (g/mol)}}$$

$$= \frac{34.02\ \cancel{\text{g/mol}}}{17.01\ \cancel{\text{g/mol}}} = 2$$

The molecular formula is $\mathbf{H_2O_2}$.

Skill 3-5 : Lactic acid ($\mathcal{M}$ = 90.08 g/mol) contains 40.0 mass % C, 6.71 mass % H, and 53.3 mass % O.

(a) Determine the empirical formula of lactic acid.

(b) Determine the molecular formula.

Plan: Assume 100 g lactic acid to express each mass % as grams. Convert grams to moles and find the empirical formula.

Solution: Express mass % as grams; assume 100 g lactic acid:

$$\text{Mass (g) of C} = \frac{40.0 \text{ parts C by mass}}{100 \text{ parts by mass}} \times 100 \text{ g} = 40.0 \text{ g C}$$

Similarly, there are 6.71 g H and 53.3 g O.

Converting from grams of elements to moles:

$$\text{Moles of C} = \text{mass of C} \times \frac{1}{\mathcal{M} \text{ of C}} = 40.0 \cancel{\text{g C}} \times \frac{1 \text{ mol C}}{12.01 \cancel{\text{g C}}}$$

$$= 3.33 \text{ mol C}$$

Similarly, we have 6.66 mol H and 3.33 mol O.

Constructing the preliminary formula: $C_{3.33}H_{6.66}O_{3.33}$

Converting to integer subscripts:

$$C_{\frac{3.33}{3.33}}H_{\frac{6.66}{3.33}}O_{\frac{3.33}{3.33}} \rightarrow C_1H_2O_1$$; the empirical formula is $\mathbf{CH_2O}$

Determining the molecular formula:

Plan: Divide the molar mass by the empirical formula mass.

Solution:

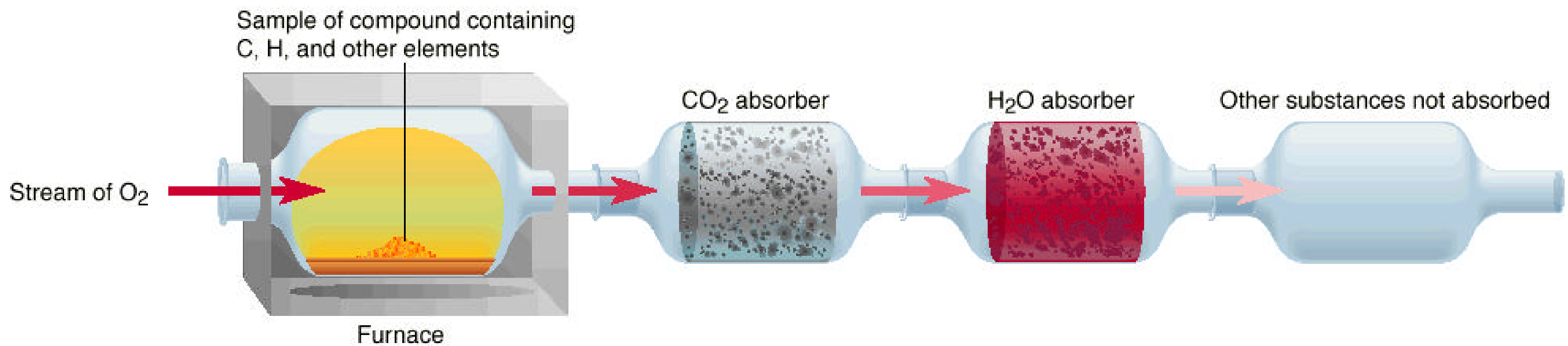
$$\text{Whole-number multiple} = \frac{\mathcal{M}\text{ of lactic acid}}{\text{empirical formula mass}} = \frac{90.08\ \cancel{\text{g/mol}}}{30.03\ \cancel{\text{g/mol}}} = 3.000 = 3$$

Calculate the molecular formula:

$$C_{(1\times 3)}H_{(2\times 3)}O_{(1\times 3)} = \mathbf{C_3H_6O_3}$$

Combustion Analysis

- Combustion analysis gives the amounts of carbon and hydrogen in a sample of combustible compound.
- The C forms CO_2, which is absorbed in the first chamber.
- The H forms H_2O, which is absorbed in the second chamber.



- Weigh the chambers before and after combustion.
- Calculate the masses of CO_2 and H_2O
- Use them to calculate the masses of C and H in the compound.
- From the masses, calculate the empirical formula.

Skill 3-5 contd: Vitamin C (M= 176. g/mol) contains C, H and O. A 1.000 g sample was placed in a combustion apparatus:

Mass of CO_2 absorber after combustion = 85.35 g

Mass of CO_2 absorber before combustion = 83.85 g

Mass of H_2O absorber after combustion = 37.96 g

Mass of H_2O absorber before combustion = 37.55 g

What is the molecular formula of vitamin C?

Plan: -Use changes in mass of CO_2 and H_2O absorbers to calculate the moles of C and H present in the sample.

-Find the mass of C, using the mass fraction of C in CO_2. Likewise, find the mass of H from the mass of H_2O.

-The mass of vitamin C minus the sum of the C and H masses gives the mass of O.

-Construct the empirical formula and molecular formula.

Solution: Finding the masses of combustion products:

$$\text{Mass of } CO_2 = \text{mass of } CO_2 \text{ absorber after} - \text{mass before} = 1.50 \text{ g } CO_2$$

$$\text{Mass of } H_2O = \text{mass of } H_2O \text{ absorber after} - \text{mass before} = 0.41 \text{ g } H_2O$$

Calculating mass fractions of the elements:

$$\text{Mass fraction of C in } CO_2 = \frac{\text{mol C} \times \mathcal{M} \text{ of C}}{\text{mass of 1 mol } CO_2} = \frac{1 \cancel{\text{mol C}} \times \dfrac{12.01 \text{ g C}}{1 \cancel{\text{mol C}}}}{44.01 \text{ g } CO_2}$$

$$= 0.2729 \text{ g C/1 g } CO_2$$

$$\text{Mass fraction of H in } H_2O = \frac{\text{mol H} \times \mathcal{M} \text{ of H}}{\text{mass of 1 mol } H_2O} = \frac{2 \cancel{\text{mol H}} \times \dfrac{1.008 \text{ g H}}{1 \cancel{\text{mol H}}}}{18.02 \text{ g } H_2O}$$

$$= 0.1119 \text{ g H/1 g } H_2O$$

Calculate masses of C and H:

$$\text{Mass of element} = \text{mass of compound} \times \text{mass fraction of element}$$

$$\text{Mass (g) of C} = 1.50\ \cancel{\text{g } CO_2} \times \frac{0.2729\text{ g C}}{1\ \cancel{\text{g } CO_2}} = 0.409\text{ g C}$$

$$\text{Mass (g) of H} = 0.41\ \cancel{\text{g } H_2O} \times \frac{0.1119\text{ g H}}{1\ \cancel{\text{g } H_2O}} = 0.046\text{ g H}$$

Calculate mass of O by difference:

$$\text{Mass (g) of O} = \text{mass of vitamin C sample} - (\text{mass of C} + \text{mass of H})$$

$$= 1.000\text{ g} - (0.409\text{ g} + 0.046\text{ g}) = 0.545\text{ g O}$$

Find the moles of elements: Divide mass of each by its molar mass gives 0.0341 mol C, 0.046 mol H, and 0.0341 mol O.

Construct the preliminary formula: $C_{0.0341}H_{0.046}O_{0.0341}$

Divide through by the smallest subscript:

$$C_{\frac{0.0341}{0.0341}}H_{\frac{0.046}{0.0341}}O_{\frac{0.0341}{0.0341}} = C_{1.00}H_{1.3}O_{1.00}$$

Determine the empirical formula. By inspection, find that 3 is the smallest multiple that makes all subscripts into integers:

$$C_{(1.00\times 3)}H_{(1.3\times 3)}O_{(1.00\times 3)} = C_{3.00}H_{3.9}O_{3.00} \approx C_3H_4O_3$$

Determine the molecular formula:

$$\text{Whole-number multiple} = \frac{\mathcal{M}\text{ of vitamin C}}{\text{empirical formula mass}} = \frac{176.12\ \cancel{\text{g/mol}}}{88.06\ \cancel{\text{g/mol}}}$$

$$= 2.000 = 2$$

$$C_{(3\times 2)}H_{(4\times 2)}O_{(3\times 2)} = \mathbf{C_6H_8O_6}$$

Chemical Formulas and the Structures of Molecules

- Chemical formulas represent real three-dimensional objects.

- The molecular formula tells the number of each type of atom.

- Molecular formulas may not be unique; the same types and numbers of atoms can bond to each other in more than one way

- Can be different compounds or structural isomers of the same compound (same molecular formulas).

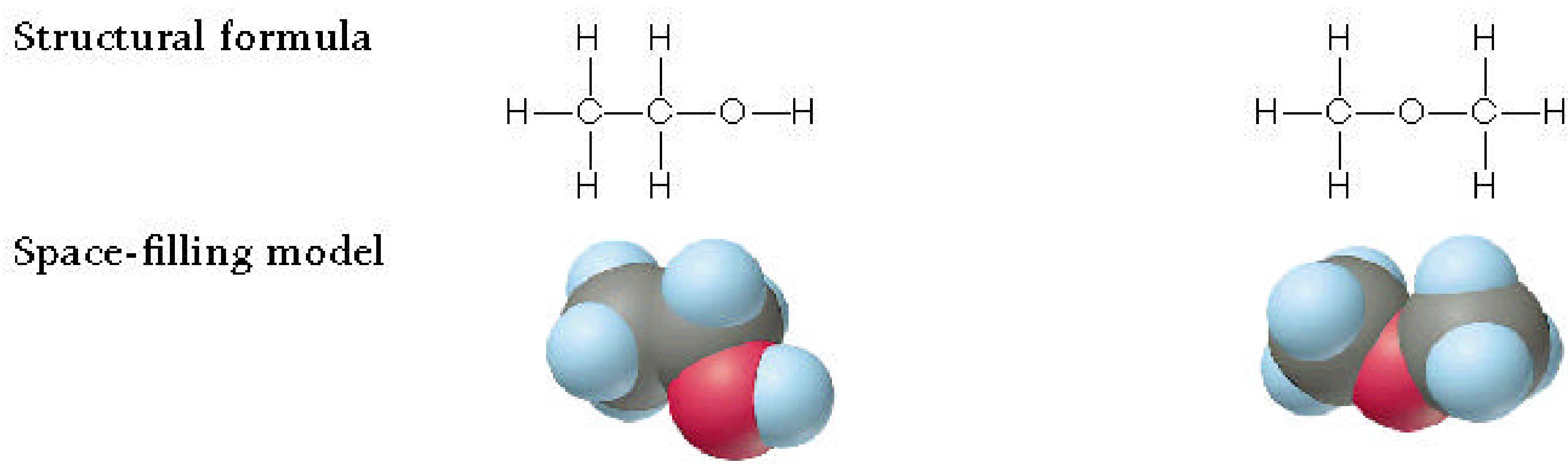
Structural Formulas

• Consider C_2H_6O; two very different compounds have this molecular formula:

-ethanol, the intoxicating substance present in wine and beer
-dimethyl ether, a colorless gas once used in refrigeration

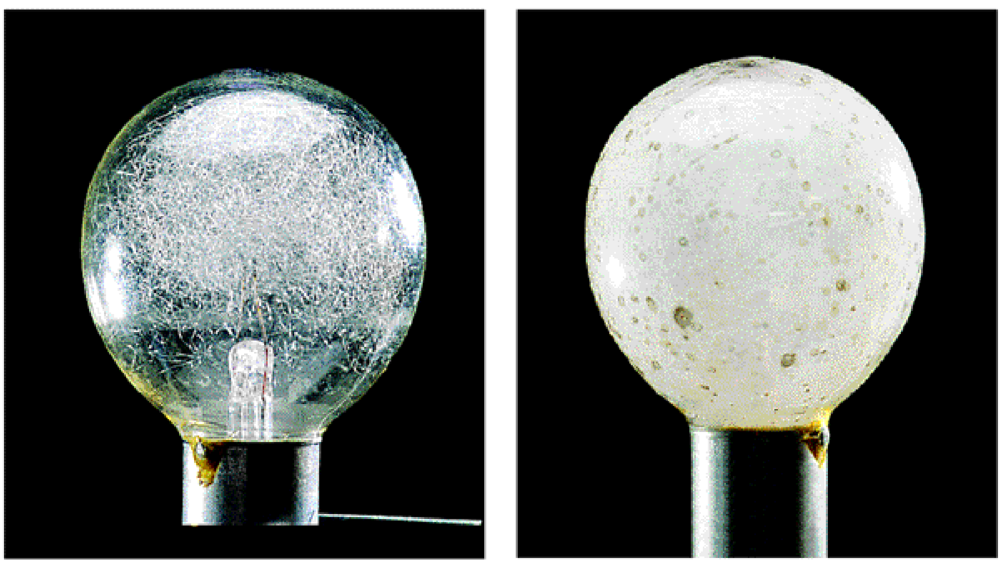
• Their radically different physical and chemical behaviors are the result of different molecular structures.

ethanol   dimethyl ether



Concept 6. The importance of balancing equations for the quantitative study of chemical reactions

- For an equation to depict accurately the amounts of chemicals involved in a reaction, it must be balanced.
- The same number of each type of atom must appear on both sides of the equation.
- In a flashbulb Mg wire and oxygen gas react to yield magnesium oxide powder.



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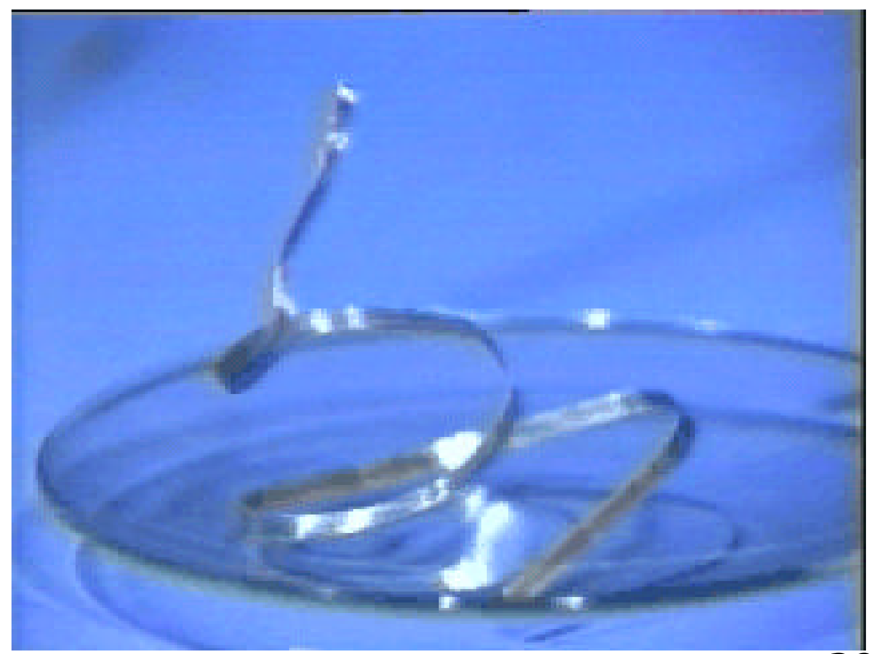


Translate the chemical statement into a "skeleton" equation: chemical formulas arranged in an equation format.

reactants — "yield" → products

$$\underset{}{\_\_\ Mg} \overset{\text{reactants}}{+} \_\_\ O_2 \xrightarrow{\text{yield}} \overset{\text{product}}{\_\_\ MgO}$$

Balancing the atoms

- Each blank must contain a balancing coefficient, a numerical multiplier of all the atoms in the formula that follows it.
- Stepwise, match the atoms on each side, element by element, beginning with the most complex substance.

$$\_\_\ Mg + \_\_\ O_2 \rightarrow \underline{\ 1}MgO$$

Requires one Mg on the left. $\underline{\ 1\ }Mg + \underline{\quad}\ O_2 \rightarrow \underline{\ 1\ }MgO$

And one oxygen on the left. $\underline{\ 1\ }Mg + \underline{\tfrac{1}{2}}O_2 \rightarrow \underline{\ 1\ }MgO$

Adjusting the coefficients.

- The smallest whole-number coefficients are required.
- To remove the 1/2 O_2, multiply the whole equation by 2:

$$2Mg + 1O_2 \rightarrow 2MgO$$

- A coefficient of 1 is implied by the presence of the substance and is not written:

$$2Mg + O_2 \rightarrow 2MgO$$

- Check the numbers of atoms:

$$\text{Reactants (2 Mg, 2 O)} \rightarrow \text{products (2 Mg, 2 O)}$$

Reaction equation is balanced.

Specifying the states of matter.

- A complete equation shows the physical state of each substance or whether it is dissolved in water.

  solid (s) liquid (l) gas (g) aqueous solution (aq)

- From the statement, Mg "wire" is solid, O_2 is a gas, and "powdery" MgO is a solid:

$$2Mg(s) + O_2(g) \longrightarrow 2MgO(s)$$

Skill 3-6: Octane (C_8H_{18}) burns to form carbon dioxide and water vapor. Write a balanced equation for this reaction.

Solution: Translate the statement into a skeleton equation (with coefficient blanks). Octane and oxygen are reactants; oxygen is molecular O_2. Carbon dioxide and water vapor are products:

$$\_\_ C_8H_{18} + \_\_ O_2 \rightarrow \_\_ CO_2 + \_\_ H_2O$$

Balance the atoms.

Start with the more complex reactant, C_8H_{18}, then balance O_2:

$$\underline{\ 1\ }C_8H_{18} + \_\_ O_2 \rightarrow \_\_ CO_2 + \_\_ H_2O$$

8 C on left requires 8 CO_2 on right:

$$\underline{\ 1\ }C_8H_{18} + \_\_ O_2 \rightarrow \underline{\ 8\ }CO_2 + \_\_ H_2O$$

18 H on left requires **9** H_2O.

$$\underline{\ 1\ }C_8H_{18} + \underline{\quad} O_2 \rightarrow \underline{\ 8\ }CO_2 + \underline{\ 9\ }H_2O$$

25 O on right requires **25/2** O_2

$$\underline{\ 1\ }C_8H_{18} + \underline{25/2}\ O_2 \rightarrow \underline{\ 8\ }CO_2 + \underline{\ 9\ }H_2O$$

Adjust coefficients. $2C_8H_{18} + 25O_2 \rightarrow 16CO_2 + 18H_2O$

Check that the equation is balanced:

reactants (16 C, 36 H, 50 O) = products (16 C, 36 H, 50 O)

States of matter: C_8H_{18} is liquid; O_2, CO_2, and H_2O vapor are gases:

$$\mathbf{2C_8H_{18}(\mathit{l}) + 25O_2(\mathit{g}) \rightarrow 16CO_2(\mathit{g}) + 18H_2O(\mathit{g})}$$

Concept 7 The mole-mass-number information contained in a balanced equation

- A balanced equation relates quantities of atoms, molecules, formula units, moles of substance, and masses.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

1 molecule C_3H_8 + 5 molecules $O_2 \rightarrow$ 3 molecules CO_2 + 4 molecules H_2O

1 mol C_3H_8 + 5 mol $O_2 \rightarrow$ 3 mol CO_2 + 4 mol H_2O

44.09 amu C_3H_8 + 160.00 amu $O_2 \rightarrow$ 132.03 amu CO_2 + 72.06 amu H_2O

44.09 g C_3H_8 + 160.00 g $O_2 \rightarrow$ 132.03 g CO_2 + 72.06 g H_2O

204.09 g $\rightarrow$ 204.09 g

Concept 8 The relation between amounts of reactants and product

- Quantitative relationships in a balanced equation are expressed as stoichiometrically equivalent molar ratios.
- Use stoichiometrically equivalent molar ratios to determine how much of one substance forms from (or reacts with) another.

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

- 3 mol CO_2 is stoichiometrically equivalent to 4 mol H_2O, 5 mol O_2 is stoichiometrically equivalent to 3 mol CO_2 , and so on.
- In the combustion of propane, how many moles of O_2 are consumed when 10.0 mol H_2O are produced?

starting amt. molar ratio equiv. amt.

$$\text{Moles of } O_2 \text{ consumed} = 10.0\ \cancel{\text{mol } H_2O} \times \frac{5 \text{ mol } O_2}{4\ \cancel{\text{mol } H_2O}} = 12.5 \text{ mol } O_2$$

Skill 3-7:

Copper is obtained from sulfide ores by roasting the ore with O_2 to form powdered copper(I) oxide and gaseous sulfur dioxide.

(a) How many moles of oxygen are required to roast 10.0 mol copper(I) sulfide?

Write the balanced equation first. The reactants are Cu_2S and O_2, and the products are Cu_2O and SO_2:

$$2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g)$$

Given the moles of Cu_2S; need to find the moles of O_2.

$$\text{Moles of } O_2 = 10.0 \cancel{\text{ mol } Cu_2S} \times \frac{3 \text{ mol } O_2}{2 \cancel{\text{ mol } Cu_2S}} = \mathbf{15.0\ mol\ O_2}$$

(b) How many grams of sulfur dioxide are formed when 10.0 mol copper(I) sulfide are roasted?

Find mass of SO_2 that forms from the given mol of reactant (Cu_2S).

$$\text{Mass (g) of } SO_2 = 10.0\ \cancel{\text{mol } Cu_2S} \times \frac{2\ \cancel{\text{mol } SO_2}}{2\ \cancel{\text{mol } Cu_2S}} \times \frac{64.07\ \text{g } SO_2}{1\ \cancel{\text{mol } SO_2}} = \mathbf{641\ g\ SO_2}$$

(c) How many kg of oxygen are needed to form 2.86 kg Cu_2O?

Convert mass Cu_2O to mol:

$$\text{Moles of } Cu_2O = 2.86\ \cancel{\text{kg } Cu_2O} \times \frac{10^3\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \text{mol } Cu_2O}{143.10\ \cancel{\text{g } Cu_2O}}$$

$$= 20.0\ \text{mol } Cu_2O$$

Mol Cu_2O to mol O_2: $$\text{Moles of } O_2 = 20.0\ \cancel{\text{mol } Cu_2O} \times \frac{3\ \text{mol } O_2}{2\ \cancel{\text{mol } Cu_2O}}$$

$$= 30.0\ \text{mol } O_2$$

Mol O_2 to mass O_2: $$\text{Mass (kg) of } O_2 = 30.0\ \cancel{\text{mol } O_2} \times \frac{32.00\ \cancel{\text{g}}\ O_2}{1\ \cancel{\text{mol } O_2}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}}$$

$$= \mathbf{0.960\ kg\ O_2}$$

Chemical Reactions That Occur in a Sequence

- A product of one reaction can become a reactant of the next (reaction sequence).
- If a substance forms in one reaction and is used up in the next, write an overall equation that eliminates the substance altogether.

Skill 3-8: Roasting chalcocite is the 1st step in extracting copper.

$$2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g) \quad \text{[equation 1]}$$

Next step: copper(I) oxide reacts with powdered carbon to yield copper metal and carbon monoxide gas:

$$Cu_2O(s) + C(s) \rightarrow 2Cu(s) + CO(g) \quad \text{[equation 2]}$$

(a) Write a balanced overall equation for the two-step sequence.

Add the equations together. Only Cu_2O appears as a product in one equation and a reactant in the other, so it is the common substance.

- Two mol Cu_2O forms in eqn. 1; but one mol Cu_2O reacts in eqn. 2.
- Double all the coefficients in eq. 2; the Cu_2O from eqn. 1 is used up

$$2Cu_2S(s) + 3O_2(g) \rightarrow 2Cu_2O(s) + 2SO_2(g) \text{ [equation 1]}$$

$$2Cu_2O(s) + 2C(s) \rightarrow 4Cu(s) + 2CO(g) \text{ [equation 2 doubled]}$$

$$2Cu_2S(s) + 3O_2(g) + \cancel{2Cu_2O(s)} + 2C(s) \rightarrow \cancel{2Cu_2O(s)} + 2SO_2(g) + 4Cu(s) + 2CO(g)$$

$$\mathbf{2Cu_2S(s) + 3O_2(g) + 2C(s) \rightarrow 2SO_2(g) + 4Cu(s) + 2CO(g)}$$

Check: Reactants (4Cu, 2S, 6O, 2C) = products (4Cu, 2S, 6O, 2C)

(b) Determine the mass (kg) of copper formed per metric ton of SO_2 that is produced.

Convert mass of SO_2 to moles, apply the molar ratio from the overall equation to find mol of Cu, and then convert mol to mass (kg) of Cu.

Convert from mass of SO_2 to moles of SO_2 :

$$\text{Moles of } SO_2 = 1.00\ \cancel{\text{t } SO_2} \times \frac{10^3\ \cancel{\text{kg}}}{1\ \cancel{\text{t}}} \times \frac{10^3\ \cancel{\text{g}}}{1\ \cancel{\text{kg}}} \times \frac{1\ \text{mol } SO_2}{64.07\ \cancel{\text{g } SO_2}}$$

$$= 1.56 \times 10^4\ \text{mol } SO_2$$

Convert from moles of SO_2 to moles of Cu:

$$\text{Moles of Cu} = 1.56 \times 10^4\ \cancel{\text{mol } SO_2} \times \frac{4\ \text{mol Cu}}{2\ \cancel{\text{mol } SO_2}} = 3.12 \times 10^4\ \text{mol Cu}$$

Convert from moles of Cu to mass of Cu:

$$\text{Mass (kg) of Cu} = 3.12 \times 10^4\ \cancel{\text{mol Cu}} \times \frac{63.55\ \cancel{\text{g}}\ \text{Cu}}{1\ \cancel{\text{mol Cu}}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}}$$

$$= \mathbf{1.98 \times 10^3\ kg\ Cu}$$

Concept 9 Why one reactant limits the yield of product

- If amounts of two reactants are not in the stoichiometric ratio of a reaction, one will be in excess; the other is the limiting reactant.

Problem: Rocket fuel, [hydrazine (N_2H_4) and dinitrogen tetraoxide (N_2O_4)], reacts to form N_2 gas and water vapor. How many grams of nitrogen gas form from 1.00×10^2 g N_2H_4 and 2.00×10^2 g N_2O_4?

Write the balanced equation:

$$2N_2H_4(l) + N_2O_4(l) \longrightarrow 3N_2(g) + 4H_2O(g)$$

Convert mass of reactants to mol and find the mol of N_2 each forms.

$$\text{Moles of } N_2H_4 = 1.00 \times 10^2 \cancel{\text{g } N_2H_4} \times \frac{1 \text{ mol } N_2H_4}{32.05 \cancel{\text{g } N_2H_4}} = 3.12 \text{ mol } N_2H_4$$

$$\text{Moles of } N_2 = 3.12 \cancel{\text{mol } N_2H_4} \times \frac{3 \text{ mol } N_2}{2 \cancel{\text{mol } N_2H_4}} = 4.68 \text{ mol } N_2$$

- Whichever yields less N_2 is the limiting reactant.

$$\text{Moles of } N_2O_4 = 2.00 \times 10^2 \cancel{\text{g } N_2O_4} \times \frac{1 \text{ mol } N_2O_4}{92.02 \cancel{\text{g } N_2O_4}} = 2.17 \text{ mol } N_2O_4$$

$$\text{Moles of } N_2 = 2.17 \cancel{\text{mol } N_2O_4} \times \frac{3 \text{ mol } N_2}{1 \cancel{\text{mol } N_2O_4}} = 6.51 \text{ mol } N_2$$

N_2H_4 is the limiting reactant because less N_2 can form (4.68 mol) when all the N_2H_4 reacts.

- Convert the lower number of moles of N_2 to mass of N_2.

$$\text{Mass (g) of } N_2 = 4.68 \cancel{\text{mol } N_2} \times \frac{28.02 \text{ g } N_2}{1 \cancel{\text{mol } N_2}} = \mathbf{131 \text{ g } N_2}$$

Concept 10. The causes of lower-than-expected yields and the distinction between theoretical and actual yields

- Ideally 100% of the limiting reactant becomes product.

- This is the theoretical yield, amount of product indicated by the stoichiometrically equivalent molar ratio in balanced equation.

- The theoretical yield is rarely obtained due to several factors.

- The amount of product actually obtained is the actual yield.

- The percent yield (% yield) is the actual yield expressed as a percent of the theoretical yield:

$$\%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

Problem: Silicon carbide (SiC) is made by heating sand (silicon dioxide) with carbon. Carbon monoxide is also formed. From 100.0 kg sand, 51.4 kg SiC are formed. What is the percent yield of SiC?

Plan: Given the actual yield of SiC, need its theoretical yield to be able to calculate its percent yield. **Steps**:

-Write the balanced equation, convert mass of SiO_2 to mol, find the mol of SiC from the molar ratio, and convert mol of SiC to mass to obtain the theoretical yield and the percent yield.

Balanced equation. $SiO_2(s) + 3C(s) \rightarrow SiC(s) + 2CO(g)$

Convert the mass of SiO_2 to moles:

$$\text{Moles of SiO}_2 = 100.0\,\cancel{\text{kg SiO}_2} \times \frac{1000\,\cancel{\text{g}}}{1\,\cancel{\text{kg}}} \times \frac{1\text{ mol SiO}_2}{60.09\,\cancel{\text{g SiO}_2}} = 1664\text{ mol SiO}_2$$

$$\text{Moles of SiO}_2 = \text{moles of SiC} = 1664\text{ mol SiC}$$

Converting from moles of SiC to mass:

$$\text{Mass (kg) of SiC} = 1664\ \cancel{\text{mol SiC}} \times \frac{40.10\ \text{g SiC}}{1\ \cancel{\text{mol SiC}}} \times \frac{1\ \text{kg}}{1000\ \cancel{\text{g}}} = 66.73\ \text{kg SiC}$$

Calculating the percent yield:

$$\%\ \text{yield SiC} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100 = \frac{51.4\ \cancel{\text{kg}}}{66.73\ \cancel{\text{kg}}} \times 100 = \mathbf{77.0\%}$$

- In multistep reaction sequences, the percent yield of each reaction step is multiplied over the whole sequence.
- A six-step sequence has a theoretical yield of 35.0 g of final produc If each step has a 90.0% yield, what would be over-all actual yield?

$$\text{Actual yield} = 35.0\ \text{g} \times 0.900 \times 0.900 \times 0.900 \times 0.900 \times 0.900 \times 0.900 = 18.6\ \text{g}$$

$$\%\ \text{yield} = \frac{18.6\ \text{g}}{35.0\ \text{g}} \times 100 = 53.1\%$$

PT

Concept 11. The meanings of concentration and molarity

Solution Concentration and the Calculation of Molarity

- Most reactions occur in solution, so quantitative aspects of solution reactions are important in chemistry and other sciences

- Need to know the concentration of reactants - the number of mol in a given volume - to calculate the volume of solution that contains a given number of mol.

- Solutions consist of a lesser amount the **solute**, dissolved in a larger amount of another substance, the **solvent**.

- Molarity (M) expresses solution concentration in terms of moles of solute per liter of solution:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad \text{or} \quad M = \frac{\text{mol solute}}{\text{L soln}}$$

Problem: Hydrochloric acid is a solution of HCl gas in water. Calculate the molarity of hydrochloric acid solution if 455 mL contains 1.80 mol hydrogen chloride.

Plan:

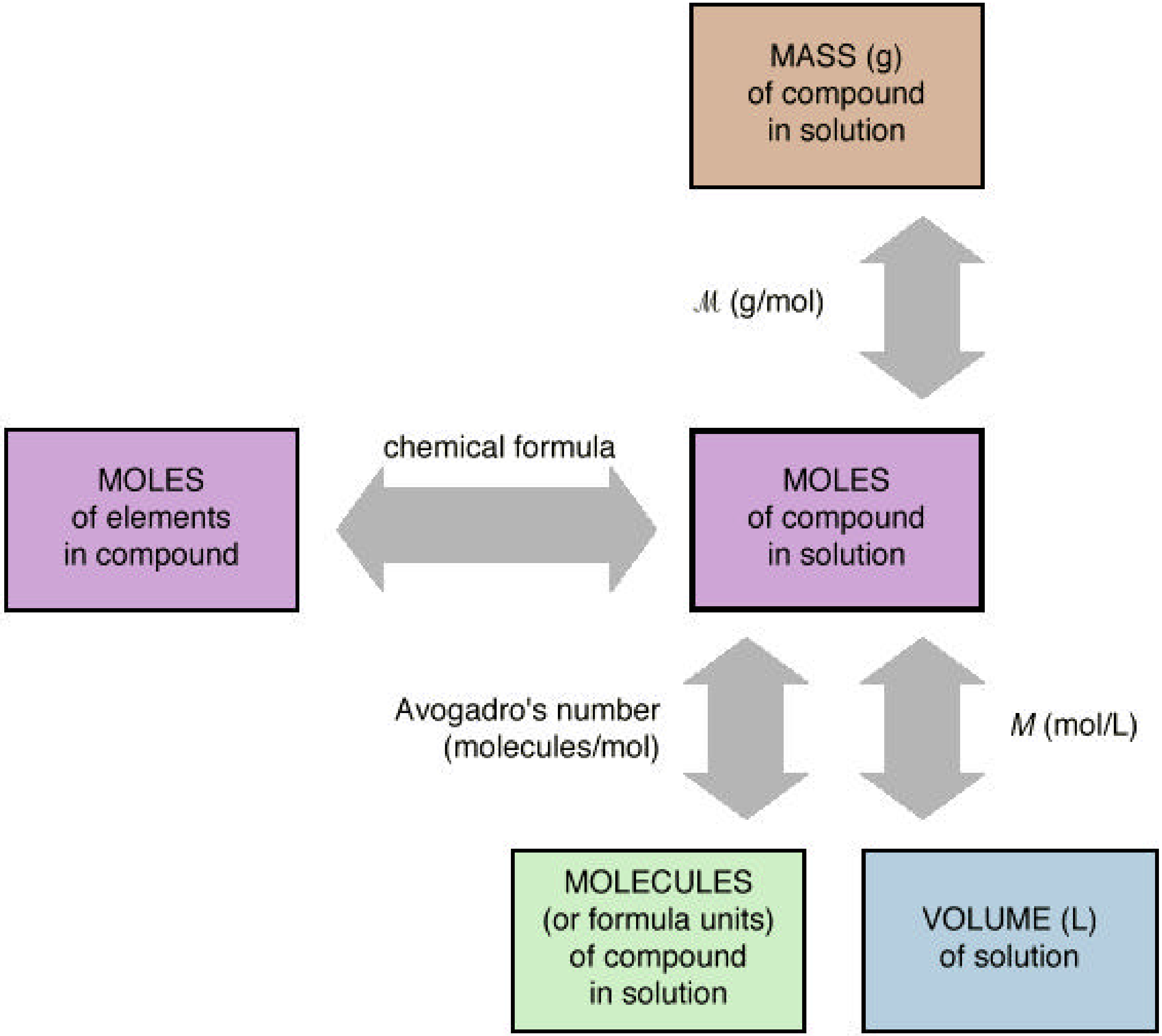
$$\text{Moles of HCl} \xrightarrow{\text{divide by volume (mL)}} \text{Concentration (mol/mL) of HCl} \xrightarrow{10^3\ \text{mL} = 1\ \text{L}} \text{Molarity (mol/L) of HCl}$$

Solution:

$$\text{Molarity} = \frac{1.80\ \text{mol HCl}}{455\ \cancel{\text{mL soln}}} \times \frac{1000\ \cancel{\text{mL}}}{1\ \text{L}} = \mathbf{3.96}\ \boldsymbol{M}\ \mathbf{HCl}$$

- Molarity is a conversion factor that extends the stoichiometric relationships among substances.

- It can be used to convert between volume of solution and moles of solute.


MASS (g) of compound in solution
$\mathcal{M}$ (g/mol)
MOLES of elements in compound
chemical formula
MOLES of compound in solution
Avogadro's number (molecules/mol)
M (mol/L)
MOLECULES (or formula units) of compound in solution
VOLUME (L) of solution

Problem: How many grams of solute are in 1.75 L of 0.460 M sodium monohydrogen phosphate?

Plan:

$$\text{Volume (L) of solution} \xrightarrow{M\ (\text{mol/L})} \text{Moles of solute} \xrightarrow{\mathcal{M}\ (\text{g/mol})} \text{Mass (g) of solute}$$

Solution: Calculating moles of solute in solution:

$$\text{Moles of } Na_2HPO_4 = 1.75\ \cancel{\text{L soln}} \times \frac{0.460 \text{ mol } Na_2HPO_4}{1\ \cancel{\text{L soln}}}$$

$$= 0.805 \text{ mol } Na_2HPO_4$$

Converting from moles of solute to mass:

$$\text{Mass (g) } Na_2HPO_4 = 0.805\ \cancel{\text{mol } Na_2HPO_4} \times \frac{141.96 \text{ g } Na_2HPO_4}{1\ \cancel{\text{mol } Na_2HPO_4}}$$

$$= \mathbf{114\ g\ Na_2HPO_4}$$

Laboratory Preparation of Molar Solutions

- Volume term in molarity is solution volume, not solvent volume.
- Solution volume includes contributions from both solute and solvent.
- The solute and solvent volumes are not additive, so measure solution volume.

To prepare 0.500 L of 0.350 M nickel(II) nitrate hexahydrate, [$Ni(NO_3)_2 \cdot 6H_2O$]:

1. Weigh out 50.9 g of the solid.

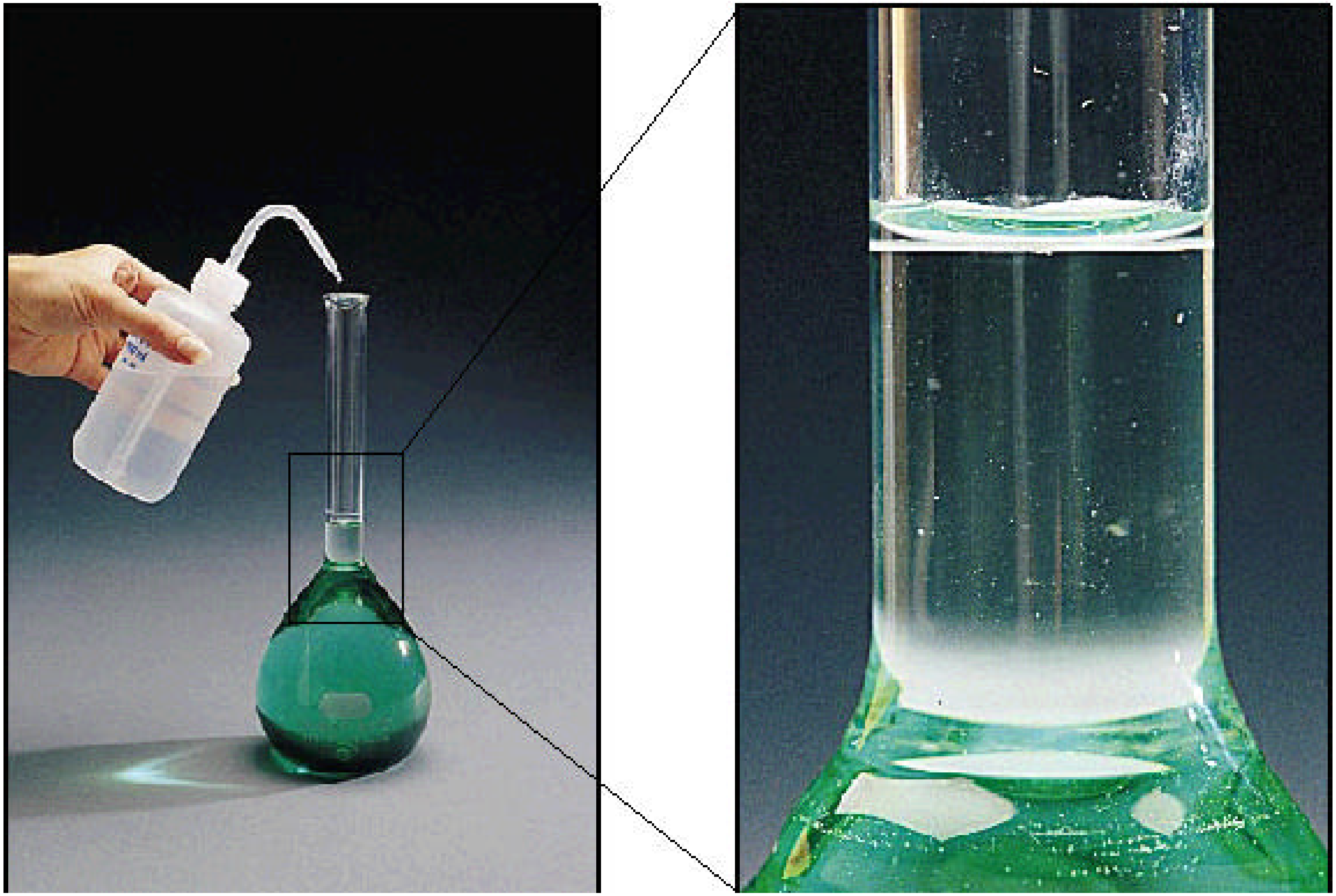
$$\text{Mass (g) of solute} = 0.500\ \cancel{\text{L soln}} \times \frac{0.350\ \cancel{\text{mol Ni(NO}_3)_2\cdot 6\text{H}_2\text{O}}}{1\ \cancel{\text{L soln}}}$$

$$\times \frac{290.83\ \text{g Ni(NO}_3)_2\cdot 6\text{H}_2\text{O}}{1\ \cancel{\text{mol Ni(NO}_3)_2\cdot 6\text{H}_2\text{O}}} = 50.9\ \text{g Ni(NO}_3)_2\cdot 6\text{H}_2\text{O}$$

2. Transfer the solid to a **volumetric flask** that contains about half the final amount of solvent.

3. Dissolve the solid completely by swirling the contents of the flask.

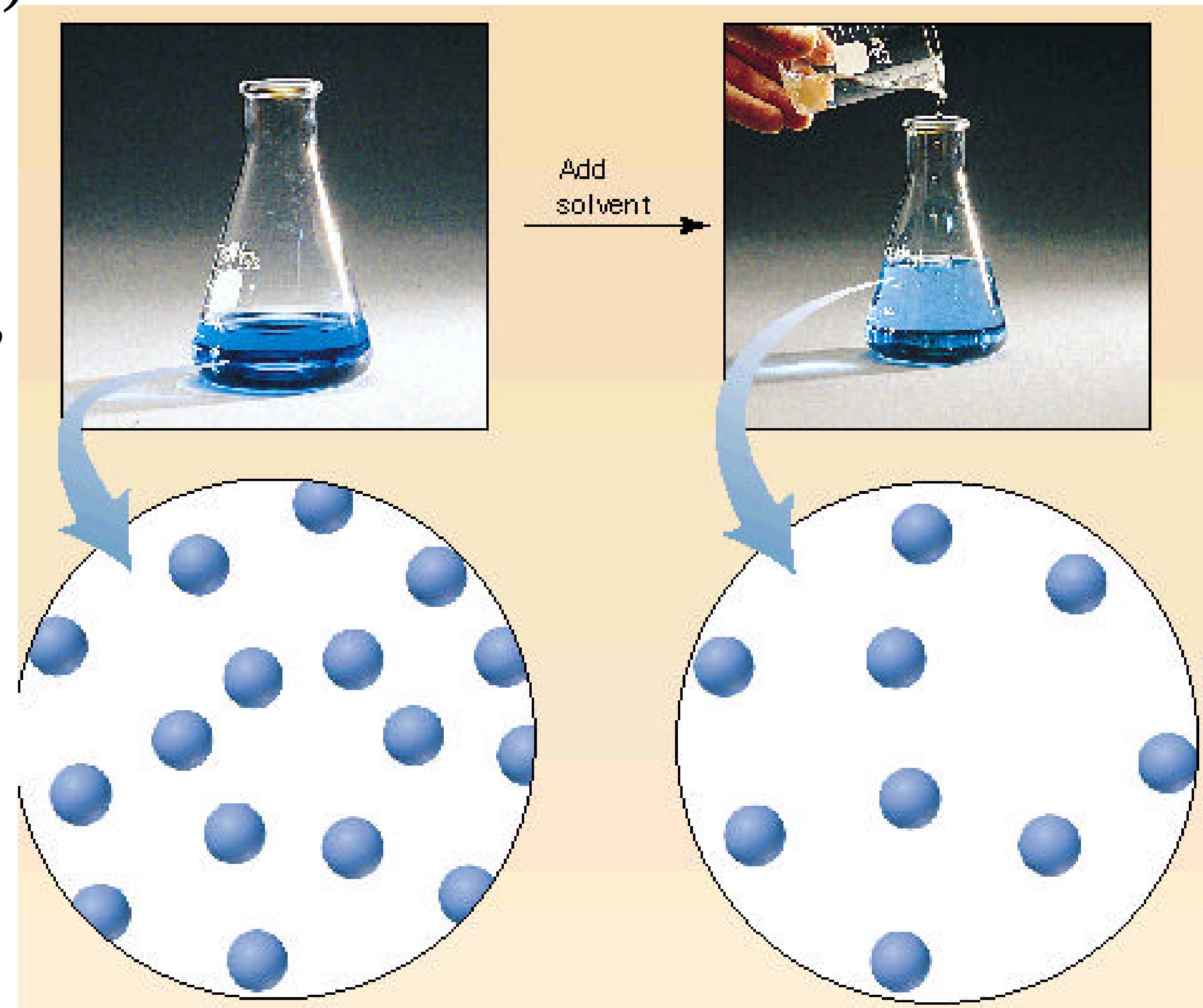
4. Add solvent until the solution reaches its final volume.



Concept 12. Effect of dilution on concentration of solute

- When sovent is added to a solution, it becomes more dilute (lower concentration).

- When lower concentrations of a solution are needed, use a more - concentrated stock solution which can be stored and diluted as needed.



Problem: Isotonic saline is 0.15 M NaCl(aq). How would you prepare 0.800 L isotonic saline from a 6.0 M stock solution?

Plan:

Volume (L) of dilute solution $\xrightarrow{M\text{ (mol/L) of dilute solution}}$ Moles of NaCl in dilute solution = moles of NaCl in concentrated solution $\xrightarrow{M\text{ (mol/L) of concentrated solution}}$ Volume (L) of concentrated solution

Solution: Finding moles of solute in dilute solution:

$$\text{Moles of NaCl in dil soln} = 0.800\ \cancel{\text{L soln}} \times \frac{0.15\text{ mol NaCl}}{1\ \cancel{\text{L soln}}} = 0.12\text{ mol NaCl}$$

Since we add only solvent to dilute the solution,

Mol of NaCl in diluted soln = mol of NaCl in concentrated sol'n.

= 0.12 mol NaCl

Finding the volume of stock solution that contains 0.12 mol NaCl:

$$\text{Volume (L) of conc NaCl soln} = 0.12\ \cancel{\text{mol NaCl}} \times \frac{1\ \text{L soln}}{6.0\ \cancel{\text{mol NaCl}}}$$

$$= 0.020\ \text{L soln}$$

To prepare 0.800 L dilute solution, place 0.020 L of 6.0 M NaCl in a flask, add distilled water to the 0.800-L mark, and stir thoroughly.

- An alternative approach to solving dilution problems makes use of the formula:

$$M_{\text{dil}} V_{\text{dil}} = \text{moles} = M_{\text{conc}} V_{\text{conc}}$$

$$V_{\text{conc}} = \frac{M_{\text{dil}} \times V_{\text{dil}}}{M_{\text{conc}}} = \frac{0.15\ \cancel{M} \times 0.800\ \text{L}}{6.0\ \cancel{M}} = 0.020\ \text{L}$$

Concept 13. How reactions in solution differ from those of pure reactants.

- Stoichiometry problems in solution need the added step of converting solution volume to mol of reactant/product .

Problem: Milk of Magnesia antacid reacts with acid to form water and magnesium chloride solution. How many liters of 0.1 M HCl react with a tablet containing 0.10 g magnesium hydroxide?

Plan: Knowing the mass of $Mg(OH)_2$ that reacts with the acid and the acid concentration, and find the acid volume.

Solution: Writing the balanced equation:

$$Mg(OH)_2(s) + 2HCl(aq) \longrightarrow MgCl_2(aq) + 2H_2O(l)$$

Converting from mass of $Mg(OH)_2$ to moles:

$$\text{Moles of } Mg(OH)_2 = 0.10\ \cancel{\text{g } Mg(OH)_2} \times \frac{1 \text{ mol } Mg(OH)_2}{58.32\ \cancel{\text{g } Mg(OH)_2}}$$

$$= 1.7 \times 10^{-3} \text{ mol } Mg(OH)_2$$

Converting from moles of $Mg(OH)_2$ to moles of HCl:

$$\text{Moles of HCl} = 1.7 \times 10^{-3}\ \cancel{\text{mol } Mg(OH)_2} \times \frac{2 \text{ mol HCl}}{1\ \cancel{\text{mol } Mg(OH)_2}} = 3.4 \times 10^{-3} \text{ mol HCl}$$

Converting from moles of HCl to volume:

$$\text{Volume (L) of HCl} = 3.4 \times 10^{-3}\ \cancel{\text{mol HCl}} \times \frac{1 \text{ L}}{0.10\ \cancel{\text{mol HCl}}} = \mathbf{3.4 \times 10^{-2}\ L}$$

Solving Limiting-Reactant Problems for Reactions in Solution

Problem: Toxic mercury compounds can be removed from solution by reaction with Na_2S solution giving solid HgS.

If 0.050 L of 0.010 M mercury(II) nitrate reacts with 0.020 L of 0.10 M sodium sulfide, how many grams of HgS can be formed?

Plan:

Balance the equation, determine the limiting reactant and convert moles of HgS it produces to mass using the molar mass of HgS.

Solution: Write the balanced equation:

$$Hg(NO_3)_2(aq) + Na_2S(aq) \longrightarrow HgS(s) + 2NaNO_3(aq)$$

Finding moles of HgS assuming $Hg(NO_3)_2$ is limiting:

$$\text{Moles of HgS} = 0.050\ \cancel{\text{L soln}} \times \frac{0.010\ \cancel{\text{mol Hg(NO}_3)_2}}{1\ \cancel{\text{L soln}}} \times \frac{1\ \text{mol HgS}}{1\ \cancel{\text{mol Hg(NO}_3)_2}}$$

$$= 5.0 \times 10^{-4}\ \text{mol HgS}$$

Finding moles of HgS assuming Na_2S is limiting:

$$\text{Moles of HgS} = 0.020\ \cancel{\text{L soln}} \times \frac{0.10\ \cancel{\text{mol Na}_2\text{S}}}{1\ \cancel{\text{L soln}}} \times \frac{1\ \text{mol HgS}}{1\ \cancel{\text{mol Na}_2\text{S}}}$$

$$= 2.0 \times 10^{-3}\ \text{mol HgS}$$

Larger than 5.0×10^{-4}, therefore, $Hg(NO_3)_2$ is the limiting reactant.

Converting from moles of HgS to mass:

$$\text{Mass (g) of HgS} = 5.0 \times 10^{-4}\ \cancel{\text{mol HgS}} \times \frac{232.7\ \text{g HgS}}{1\ \cancel{\text{mol HgS}}} = \mathbf{0.12\ g\ HgS}$$

MAIN–GROUP ELEMENTS | TRANSITION ELEMENTS | MAIN–GROUP ELEMENTS

Period	1A (1)	2A (2)	3B (3)	4B (4)	5B (5)	6B (6)	7B (7)	8B (8)	8B (9)	8B (10)	1B (11)	2B (12)	3A (13)	4A (14)	5A (15)	6A (16)	7A (17)	8A (18)
1	1 **H** 1.008																	2 **He** 4.003
2	3 **Li** 6.941	4 **Be** 9.012											5 **B** 10.81	6 **C** 12.01	7 **N** 14.01	8 **O** 16.00	9 **F** 19.00	10 **Ne** 20.18
3	11 **Na** 22.99	12 **Mg** 24.31											13 **Al** 26.98	14 **Si** 28.09	15 **P** 30.97	16 **S** 32.07	17 **Cl** 35.45	18 **Ar** 39.95
4	19 **K** 39.10	20 **Ca** 40.08	21 **Sc** 44.96	22 **Ti** 47.90	23 **V** 50.94	24 **Cr** 52.00	25 **Mn** 54.94	26 **Fe** 55.85	27 **Co** 58.93	28 **Ni** 58.70	29 **Cu** 63.55	30 **Zn** 65.39	31 **Ga** 69.72	32 **Ge** 72.59	33 **As** 74.92	34 **Se** 78.96	35 **Br** 79.90	36 **Kr** 83.80
5	37 **Rb** 85.47	38 **Sr** 87.62	39 **Y** 88.91	40 **Zr** 91.22	41 **Nb** 92.91	42 **Mo** 95.94	43 **Tc** (98)	44 **Ru** 101.1	45 **Rh** 102.9	46 **Pd** 106.4	47 **Ag** 107.9	48 **Cd** 112.4	49 **In** 114.8	50 **Sn** 118.7	51 **Sb** 121.8	52 **Te** 127.6	53 **I** 126.9	54 **Xe** 131.3
6	55 **Cs** 132.9	56 **Ba** 137.3	57 **La** 138.9	72 **Hf** 178.5	73 **Ta** 180.9	74 **W** 183.9	75 **Re** 186.2	76 **Os** 190.2	77 **Ir** 192.2	78 **Pt** 195.1	79 **Au** 197.0	80 **Hg** 200.6	81 **Tl** 204.4	82 **Pb** 207.2	83 **Bi** 209.0	84 **Po** (209)	85 **At** (210)	86 **Rn** (222)
7	87 **Fr** (223)	88 **Ra** (226)	89 **Ac** (227)	104 **Rf** (261)	105 **Ha** (262)	106 **Sg** (266)	107 **Ns** (262)	108 **Hs** (265)	109 **Mt** (266)	110	111							

INNER TRANSITION ELEMENTS

Period															
6	Lanthanides	58 **Ce** 140.1	59 **Pr** 140.9	60 **Nd** 144.2	61 **Pm** (145)	62 **Sm** 150.4	63 **Eu** 152.0	64 **Gd** 157.3	65 **Tb** 158.9	66 **Dy** 162.5	67 **Ho** 164.9	68 **Er** 167.3	69 **Tm** 168.9	70 **Yb** 173.0	71 **Lu** 175.0
7	Actinides	90 **Th** 232.0	91 **Pa** (231)	92 **U** 238.0	93 **Np** (244)	94 **Pu** (242)	95 **Am** (243)	96 **Cm** (247)	97 **Bk** (247)	98 **Cf** (251)	99 **Es** (252)	100 **Fm** (257)	101 **Md** (258)	102 **No** (259)	103 **Lr** (260)