The McGraw-Hill Companies

Unit IV - Lecture 11

Chemistry

The Molecular Nature of Matter and Change

Fifth Edition

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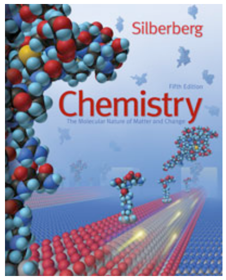


Mole - Mass Relationships in Chemical Systems

The Mole

mole(mol) - the amount of a substance that contains the same number of entities as there are atoms in exactly 12 g of carbon-12.

This amount is $\mathbf{6.022 \times 10^{23}}$. The number is called Avogadro's number and is abbreviated as N.

One mole (1 mol) contains 6.022×10^{23} entities (to four significant figures)

Figure 3.1 **Counting objects of fixed relative mass.**

A

12 red marbles @ 7g each = 84g
12 yellow marbles @ 4g each = 48g

B

55.85 g Fe = 6.022×10^{23} atoms Fe
32.07 g S = 6.022×10^{23} atoms S

Figure 3.2

One mole of common substances.

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 (or formula) 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 1 mole of chemical entities (atoms, ions, molecules, formula units)	g/mol

Table 3.2 **Information Contained in the Chemical Formula of Glucose $C_6H_{12}O_6$ (M = 180.16 g/mol)**

	Carbon (C)	Hydrogen (H)	Oxygen (O)
Atoms/molecule of compound	6 atoms	12 atoms	6 atoms
Moles of atoms/ mole of compound	6 mol of atoms	12 mol of atoms	6 mol of atoms
Atoms/mole of compound	$6(6.022 \times 10^{23})$ atoms	$12(6.022 \times 10^{23})$ atoms	$6(6.022 \times 10^{23})$ atoms
Mass/molecule of compound	6(12.01 amu) =72.06 amu	12(1.008 amu) =12.10 amu	6(16.00 amu) =96.00 amu
Mass/mole of compound	72.06 g	12.10 g	96.00 g

Interconverting Moles, Mass, and Number of Chemical Entities

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

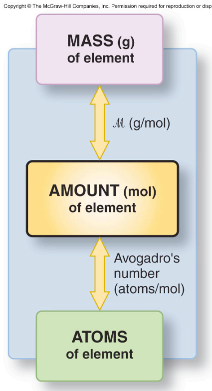
$$\text{No. of moles} = \text{mass (g)} \times \frac{\text{1 mol}}{\text{no. of grams}} \leftarrow \boxed{M}$$

$$\text{No. of entities} = \text{no. of moles} \times \frac{6.022\times10^{23}\ \text{entities}}{\text{1 mol}}$$

$$\text{No. of moles} = \text{no. of entities} \times \frac{\text{1 mol}}{6.022\times10^{23}\ \text{entities}}$$

Figure 3.3

Summary of the mass-mole-number relationships for elements.



Sample Problem 3.1 **Calculating the Mass in a Given Number of Moles of an Element**

PROBLEM: Silver (Ag) is used in jewelry and tableware but no longer in U.S. coins. How many grams of Ag are in 0.0342 mol of Ag?

PLAN: To convert mol of Ag to g, use the # g Ag/mol Ag, the molar mass ***M***.

amount (mol) of Ag

⇩ multiply by ***M*** of Ag (107.9 g/mol)

mass (g) of Ag

SOLUTION:

$$0.0342\ \text{mol Ag} \times \frac{107.9\ \text{g Ag}}{\text{mol Ag}} = 3.69\ \text{g Ag}$$

Sample Problem 3.2 **Calculating Number of Atoms in a Given Mass of an Element**

PROBLEM: Iron (Fe), the main component of steel, is the most important metal in industrial society. How many Fe atoms are in 95.8 g of Fe?

PLAN: To convert g of Fe to atoms, first find the # mols of Fe and then convert mols to atoms.

mass (g) of Fe

⇩ divide by ***M*** of Fe (55.85 g/mol)

amount (mol) of Fe

⇩ multiply by 6.022×10^{23} atoms/mol

atoms of Fe

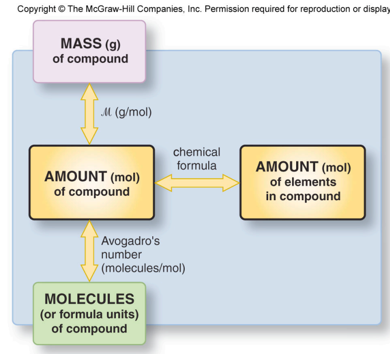
SOLUTION:

$$95.8\ \text{g Fe} \times \frac{\text{mol Fe}}{55.85\ \text{g Fe}} = 1.72\ \text{mol Fe}$$

$$1.72\ \text{mol Fe} \times \frac{6.022\times10^{23}\ \text{atoms Fe}}{\text{mol Fe}} = 1.04 \times 10^{24}\ \text{atoms Fe}$$

Figure 3.4

Summary of the mass-mole-number relationships for compounds.



Sample Problem 3.3 **Calculating the Moles and Number of Formula Units in a Given Mass of a Compound**

PROBLEM: How many formula units are in 41.6 g of ammonium carbonate?

PLAN: Write the formula for the compound. Calculate $\mathcal{M}$. Convert the given mass to mols and then mols to formula units.

mass (g) of $(NH_4)_2CO_3$

↓ divide by $\mathcal{M}$

amount (mol) of $(NH_4)_2CO_3$

↓ multiply by 6.022×10^{23} formula units/mol

number of $(NH_4)_2CO_3$ formula units

SOLUTION: The formula is $(NH_4)_2CO_3$.

$\mathcal{M}$ = (2 x 14.01 g/mol N) + (8 x 1.008 g/mol H) + (12.01 g/mol C) + (3 x 16.00 g/mol O) = 96.09 g/mol

$$41.6 \text{ g } (NH_4)_2CO_3 \times \frac{\text{mol } (NH_4)_2CO_3}{96.09 \text{ g } (NH_4)_2CO_3} \times \frac{6.022\times10^{23} \text{ formula units } (NH_4)_2CO_3}{\text{mol } (NH_4)_2CO_3} =$$

2.61×10^{23} formula units $(NH_4)_2CO_3$

Mass Percent from the Chemical Formula

Mass % of element X =

$$\frac{\text{atoms of X in formula x atomic mass of X (amu)}}{\text{molecular (or formula) mass of compound (amu)}} \text{ x 100}$$

Mass % of element X =

$$\frac{\text{moles of X in formula x molar mass of X (g/mol)}}{\text{mass (g) of 1 mol of compound}} \text{ x 100}$$

Sample Problem 3.4 **Calculating the Mass Percents and Masses of Elements in a Sample of Compound**

PROBLEM: Glucose ($C_6H_{12}O_6$) generates chemical potential energy.

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

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

PLAN: Find the total mass of each element and of the glucose molecule. Divide the mass of each element by the mass of the molecule and multiply by 100.

amount (mol) of element X in 1 mol compound

↓ multiply by $\mathcal{M}$ (g/mol) of X

mass (g) of X in 1 mol of compound

↓ divide by mass (g) of 1 mol of compound

mass fraction of X

↓ multiply by 100

mass % X in compound

Sample Problem 3.4 **Calculating the Mass Percents and Masses of Elements in a Sample of Compound**

SOLUTION: **(a)** Per mole glucose there are:

6 moles of C, **12** moles H, and **6** moles O

$6 \text{ mol C} \times \frac{12.01 \text{ g C}}{\text{mol C}} = 72.06 \text{ g C}$ $\quad$ $12 \text{ mol H} \times \frac{1.008 \text{ g H}}{\text{mol H}} = 12.096 \text{ g H}$

$6 \text{ mol O} \times \frac{16.00 \text{ g O}}{\text{mol O}} = 96.00 \text{ g O}$ $\quad$ $\mathcal{M}$ = 180.16 g/mol

(b)

mass percent of C = $\frac{72.06 \text{ g C}}{180.16 \text{ g glucose}}$ = 0.3999 x 100 = 39.99 mass %C

mass percent of H = $\frac{12.096 \text{ g H}}{180.16 \text{ g glucose}}$ = 0.06714 x 100 = 6.714 mass %H

mass percent of O = $\frac{96.00 \text{ g O}}{180.16 \text{ g glucose}}$ = 0.5329 x 100 = 53.29 mass %O

Empirical and Molecular Formulas

Empirical Formula -

The simplest formula for a compound that agrees with the elemental analysis and gives rise to the smallest set of whole numbers of atoms.

Molecular Formula -

The formula of the compound as it exists; it may be a multiple of the empirical formula.

Sample Problem 3.5 **Determining the Empirical Formula from Masses of Elements**

PROBLEM: Elemental analysis of a sample of an ionic compound showed 2.82 g of Na, 4.35 g of Cl, and 7.83 g of O. What is the empirical formula and name of the compound?

PLAN: Find the relative number of moles of each element. Divide by the lowest mol amount to find the relative mol ratios (empirical formula).

mass (g) of each element

↓ divide by $\mathcal{M}$ (g/mol)

amount (mol) of each element

↓ use # of moles as subscripts

preliminary formula

↓ change to integer subscripts

empirical formula

SOLUTION:

$2.82 \text{ g Na} \times \frac{\text{mol Na}}{22.99 \text{ g Na}} = 0.123 \text{ mol Na}$

$4.35 \text{ g Cl} \times \frac{\text{mol Cl}}{35.45 \text{ g Cl}} = 0.123 \text{ mol Cl}$

$7.83 \text{ g O} \times \frac{\text{mol O}}{16.00 \text{ g O}} = 0.489 \text{ mol O}$

Na and Cl = $\frac{0.123}{0.123}$ = 1 and O = $\frac{0.489}{0.123}$ = 3.98

→ $Na_1\ Cl_1\ O_{3.98}$ or $NaClO_4$ $\quad$ $NaClO_4$ is sodium perchlorate.

Sample Problem 3.6 Determining a Molecular Formula from Elemental Analysis and Molar Mass

PROBLEM: Elemental analysis of lactic acid (***M*** = 90.08 g/mol) shows it 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 and find the mass of each element
⇩ divide each mass by mol mass (*M*)
amount (mol) of each element
⇩ use # mols as subscripts
preliminary formula
⇩ convert to integer subscripts
empirical formula
⇩ divide mol mass by mass of empirical formula to get a multiplier
m

Sample Problem 3.6 Determining a Molecular Formula from Elemental Analysis and Molar Mass

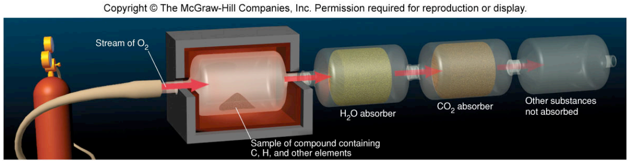
SOLUTION: Assuming there are 100. g of lactic acid, the constituents are

$40.0\text{ g C} \times \frac{\text{mol C}}{12.01\text{ g C}} = 3.33\text{ mol C}$ $6.71\text{ g H} \times \frac{\text{mol H}}{1.008\text{ g H}} = 6.66\text{ mol H}$ $53.3\text{ g O} \times \frac{\text{mol O}}{16.00\text{ g O}} = 3.33\text{ mol O}$

$C_{\frac{3.33}{3.33}}H_{\frac{6.66}{3.33}}O_{\frac{3.33}{3.33}} \longrightarrow CH_2O$ empirical formula

$\frac{\text{molar mass of lactate}}{\text{mass of } CH_2O} \longrightarrow \frac{90.08\text{ g}}{30.03\text{ g}} \longrightarrow 3$ $C_3H_6O_3$ is the molecular formula

Figure 3.5 Combustion apparatus for determining formulas of organic compounds.



$$C_nH_m + (n+\frac{m}{2})\ O_2 = n\ CO(g) + \frac{m}{2}\ H_2O(g)$$

Sample Problem 3.7 Determining a Molecular Formula from Combustion Analysis

PROBLEM: When a 1.000 g sample of vitamin C (M = 176.12 g/mol) is placed in a combustion chamber and burned, the following data are obtained:

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: difference (after - before) = mass of oxidized element
⇩
find the mass of each element in its combustion product
⇩
find the mols → preliminary formula → empirical formula → molecular formula

Sample Problem 3.7 Determining a Molecular Formula from Combustion Analysis

SOLUTION: CO_2 85.35 g - 83.85 g = 1.50 g H_2O 37.96 g - 37.55 g = 0.41 g

$1.50\text{ g } CO_2 \times \frac{12.01\text{ g C}}{44.01\text{ g } CO_2} = 0.409\text{ g C}$ $\frac{0.409\text{ g C}}{12.01\text{ g C}} = 0.0341\text{ mol C}$

$0.41\text{ g } H_2O \times \frac{2.016\text{ g H}}{18.02\text{ g } H_2O} = 0.046\text{ g H}$ $\frac{0.046\text{ g H}}{1.008\text{ g H}} = 0.0456\text{ mol H}$

1.000 g - (0.409 + 0.046) g = 0.545 g O $\frac{0.545\text{ g O}}{16.00\text{ g O}} = 0.0341\text{ mol O}$

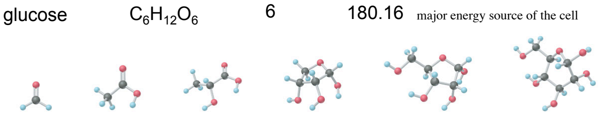
C $\frac{0.0341}{0.0341} = 1$; H $\frac{0.0456}{0.0341} = 1.34$; O $\frac{0.0341}{0.0341} = 1$

$C_1H_{1.34}O_1 = C_3H_{4.01}O_3 \longrightarrow C_3H_4O_3$

$\frac{176.12\text{ g/mol}}{88.06\text{ g}} = 2.000\text{ mol}$ $C_6H_8O_6$

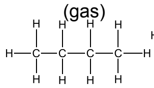
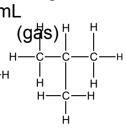
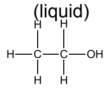
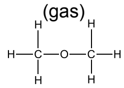
Table 3.3 Some Compounds with Empirical Formula CH_2O (Composition by Mass: 40.0% C, 6.71% H, 53.3% O)

Name	Molecular Formula	Whole-Number Multiple	*M* (g/mol)	Use or Function
formaldehyde	CH_2O	1	30.03	disinfectant; biological preservative
acetic acid	$C_2H_4O_2$	2	60.05	acetate polymers; vinegar (5% soln)
lactic acid	$C_3H_6O_3$	3	90.09	sour milk; forms in exercising muscle
erythrose	$C_4H_8O_4$	4	120.10	part of sugar metabolism
ribose	$C_5H_{10}O_5$	5	150.13	component of nucleic acids and B_2
glucose	$C_6H_{12}O_6$	6	180.16	major energy source of the cell



CH_2O $C_2H_4O_2$ $C_3H_6O_3$ $C_4H_8O_4$ $C_5H_{10}O_5$ $C_6H_{12}O_6$

Table 3.4 Two Pairs of Constitutional Isomers

	C_4H_{10}		C_2H_6O	
Property	Butane	2-Methylpropane	Ethanol	Dimethyl Ether
$\boldsymbol{M}$(g/mol)	58.12	58.12	46.07	46.07
Boiling Point	-0.5^0C	-11.06^0C	78.5^0C	-25^0C
Density at 20^0C	0.579 g/mL (gas)	0.549 g/mL (gas)	0.789 g/mL (liquid)	0.00195 g/mL (gas)
Structural formulas				
Space-filling models