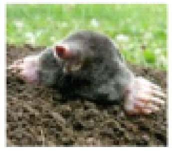


Applying the Mole Concept

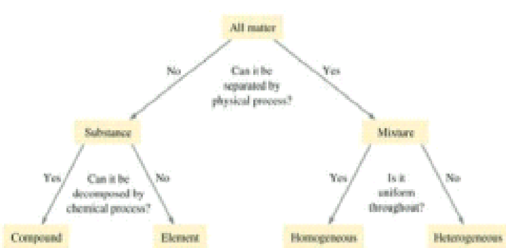
Objectives: By the end of today's class you will be able to

- calculate the molar mass of compounds
- calculate mass percent composition
- determine the number of atoms or molecules present in a given sample
- determine the empirical and molecular formula of a compound

Review of basic definitions

- *Element:* composed of *identical atoms* that cannot be decomposed into simpler substances by chemical methods.
  Ex: carbon, gold, chlorine.
- *Atom:* the smallest unit that maintains the properties of the element.
- *Isotope:* atoms of the same element having different atomic masses, i.e. same number of protons (atomic number), but different number of neutrons.
  Ex: ^{12}C, ^{13}C, ^{14}C
- *Compound:* atoms bonded together in *fixed proportions*. Compounds can be decomposed into elements by chemical methods.
  Ex: water - H_2O, sodium chloride - NaCl, acetic acid - CH_3COOH
- *Substance:* cannot be separated by physical means into portions that have different properties. Substances are formed of elements or compounds.
- *Mixture:* formed by the physical combination of two or more compounds or elements (substances). Mixtures can be homogeneous (single-phase) or heterogeneous (multi-phase).

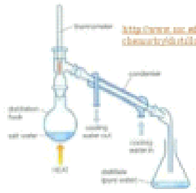
Classification of matter

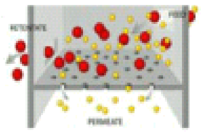


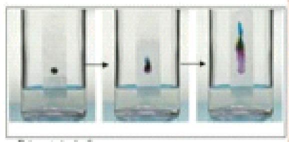
Petrucci et. al. p6

Physical separation techniques

www.eppendorf.com

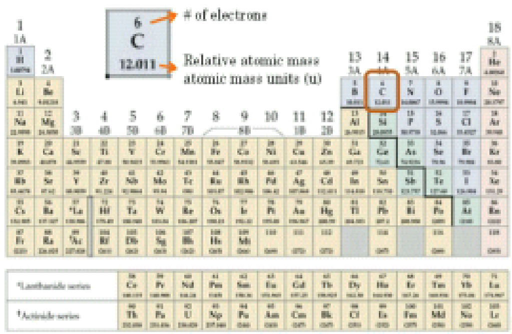






Petrucci et. al. p7

The Periodic table of Elements



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Avogadro's constant and the mole

- Let's define 1 mole of carbon as 12 g of ^{12}C (carbon-12).
- Avogadro's constant is defined as the number of elementary entities (atoms, molecules, etc.) in 1 mole.

Avogadro's constant $= \frac{12\ \text{g/mol}}{1.9926 \times 10^{-23}\ \text{g}} = 6.022 \times 10^{23}/\text{mol}$

- The mole is a measurement of quantity.

  1 mole = 6.022×10^{23} units
- Therefore,

  1 mole of O_2 molecules = 6.022×10^{23} O_2 molecules

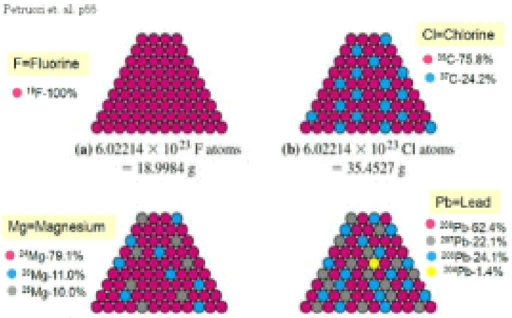
  1 mole of O atoms = 6.022×10^{23} O atoms

  1 mole of students = 6.022×10^{23} students

Molar masses

- The mass of 1 mole of a substance is called the **molar mass** or **molecular weight** of the substance.
- Since we defined 1 mole of carbon as 12 g of ^{12}C, the mass given on the periodic table is the molar mass!

Petrucci et. al. p55



(a) 6.02214×10^{23} F atoms $= 18.9984$ g

(b) 6.02214×10^{23} Cl atoms $= 35.4527$ g

(c) 6.02214×10^{23} Mg atoms $= 24.3050$ g

(d) 6.02214×10^{23} Pb atoms $= 207.2$ g

Example 1

Determine the molar mass of each of the following:

C $M_C = 12.011$ g/mol

H $M_H = 1.00794$

O $M_O = 15.9994$

$C_7H_{14}O_2$ $M_{C_7H_{14}O_2} = 130.2$

Example 2

Isopentyl acetate ($C_7H_{14}O_2$) is the compound responsible for the scent of bananas. Interestingly, bees release about 1 μg of this compound when they sting, in order to attract other bees to join the attack.

How many molecules of isopentyl acetate are released in a typical bee sting?

How many atoms of carbon are present in this sample?

Basis: 1 μg $C_7H_{14}O_2$

$M_{C_7H_{14}O_2} = 130.2$ g/mol

1 mol $= 6.022 \times 10^{23}$ molecules

$$1\,\mu g \times \frac{1\,g}{10^6\,\mu g} \times \frac{1\,mol}{130.2\,g} \times \frac{6.022 \times 10^{23}}{1\,mol}$$

#carbon = #molecules · 7

Example 3: Calculating mass percent composition

Penicillin, the first of a now large number of antibiotics, was discovered accidentally by the Scottish bacteriologist Alexander Fleming in 1928, but he was never able to isolate it as a pure compound. This and similar antibiotics have saved millions of lives that might have been lost to infection.

Penicillin F has the formula $C_{14}H_{20}N_2SO_4$. Compute the mass percent of each element.

Element	(g/mol) $M_{Element}$	(g) m	
C	12.001	168.154	53.83
H	1.00794	20.1588	6.45
N	14.0067	28.014	8.97
S	32.065	32.065	10.26
O	15.9994	63.9976	20.49
		312.4	

Empirical vs. Molecular Formulas

- The **molecular formula** shows the actual number of atoms in a molecule.
  Ex. The molecular formula of hexene is C_6H_{12}.
- The **empirical formula** gives the relative number of atoms in a molecule.
  Ex. The empirical formula of hexene is CH_2.
- The molecular formula can be obtained from the empirical formula if the molar mass of the compound is known.

Example 4:

The explosive RDX contains 16.22% carbon, 2.72% hydrogen, 37.84% nitrogen, and 43.22% oxygen by mass. The molar mass of RDX is approximately 220 g/mol.

Determine the empirical and molecular formulas of RDX.

Basis: 1 mol = 220 g RDX

Element	mass%	m(g)	M (g/mol)	n (moles)
C	16.22	35.684	12.011	2.97
H	2.72	5.984	1.00794	5.94
N	37.84	83.2	14.0067	5.94
O	43.22	95.08	15.9994	5.95

$C_3H_6N_6O_6$ Molecular

$CH_2N_2O_2$ Empirical

Example 5:

Chemical bottles labeled to contain blue-colored $NiSO_4 \cdot 6H_2O$ often contain instead the green colored $NiSO_4 \cdot 7H_2O$. How many fewer grams of Ni are in 1.0 lb of $NiSO_4 \cdot 7H_2O$ than would be in 1.0 lb of $NiSO_4 \cdot 6H_2O$?

Challenge!
Example 6:

Estimate the distance between molecules of water at 4 °C.

$$\left.\begin{matrix} H - 1.00794 \times 2 \\ O - 15.9994 \end{matrix}\right\} 18 \text{ g/mol}$$

Density = 1 g/cm^3

$$1 \text{ cm}^3 \times 1 \frac{\text{g}}{\text{cm}^3} \times \frac{1}{18.01} \frac{\text{g}}{\text{mol}} \times 6.022 \times 10^{23}$$

$= 3.34 \times 10^{22}$ molecules H_2O

Volume per $H_2O = \frac{1 \text{ cm}^3}{3.34 \times 10^{22}}$

$= 2.94 \times 10^{-22} \text{ cm}^3$

length $= \sqrt[3]{2.94 \times 10^{-22}}$

assume H_2O molecule is a cube.

Suggested readings

1.3 Classification of Matter
2.7 The Concept of the Mole and the Avogadro Constant
2.8 Using the Mole Concept in Calculations
3.1 Types of Chemical Compounds and Their Formulas
3.2 The Mole Concept and Chemical Compounds
3.3 Composition of Chemical Compounds