1.1 Moles and Molar Mass

Essential knowledge statements from the AP Chemistry CED:

- One cannot count particles directly while performing laboratory work. Thus, there must be a connection between the masses of substances reacting and the actual number of particles undergoing chemical changes.
- Avogadro's number ($N_A = 6.02\times10^{23}$ mol^{-1}) provides the connection between the number of moles in a pure sample of a substance and the number of constituent particles (or formula units) of that substance.
- Expressing the mass of an individual atom or molecule in atomic mass units (amu) is useful because the average mass in amu of one particle (atom or molecule) or formula unit of a substance will always be numerically equal to the molar mass of that substance in grams. Thus, there is a quantitative connection between the mass of a substance and the number of particles that the substance contains.

The particles of a substance can be described as atoms, molecules, or formula units, as shown in the following examples. The molar mass of a substance can be determined or calculated from the atomic mass values on the periodic table.

1 mol Mg = 24.30 g Mg = 6.02×10^{23} atoms Mg

1 mol CO_2 = 44.01 g CO_2 = 6.02×10^{23} molecules CO_2

1 mol NaCl = 58.44 g NaCl = 6.02×10^{23} formula units NaCl

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1. Calculate the mass, in grams, of 0.0850 mol $Ba(OH)_2$.

2. Calculate the number of moles of C_4H_{10} present in 2.00 g C_4H_{10}.

3. Calculate the number of atoms of Si present in 35.0 mol Si.

4. Calculate the number of moles of O_3 present in 4.3×10^{24} molecules of O_3.

5. Calculate the mass, in grams, of 8.2×10^{22} molecules of $CHCl_3$.

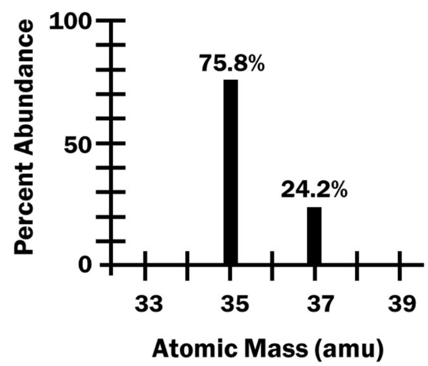
6. Calculate the number of formula units of Na_2SO_4 present in 0.248 g Na_2SO_4.

1.2 Mass Spectroscopy of Elements

Essential knowledge statements from the AP Chemistry CED:

- The mass spectrum of a sample containing a single element can be used to determine the identity of the isotopes of that element and the relative abundance of each isotope in nature.
- The average atomic mass of an element can be estimated from the weighted average of the isotopic masses using the mass of each isotope and its relative abundance.

Isotope	Abundance
Cl-35	75.8%
Cl-37	24.2%

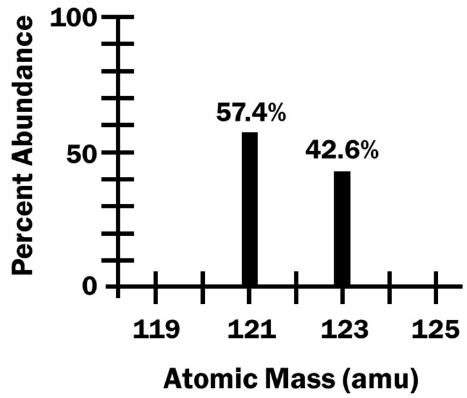


7. Based on the information shown above,

(a) calculate the average atomic mass of Cl.

(b) Fill in the table below.

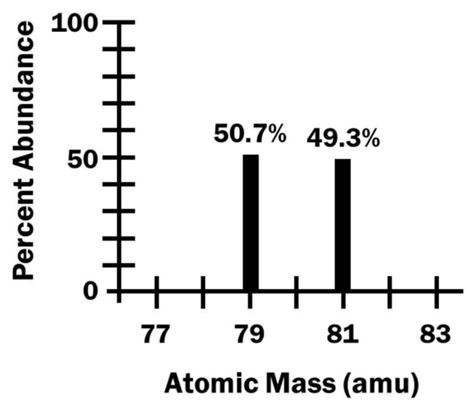
Isotope	Protons	Neutrons
Cl-35		
Cl-37		



8. Based on the information shown above,

(a) calculate the average atomic mass of the element.

(b) What is the most likely identity of this element? ____



9. Based on the information shown above,

(a) what is the most likely identity of this element? ____

(b) Fill in the table below.

Mass Number	Protons	Neutrons
79		
81		

10. A certain element has two naturally occurring isotopes with mass numbers of 63 and 65.

(a) What is the most likely identity of this element? ______________________________

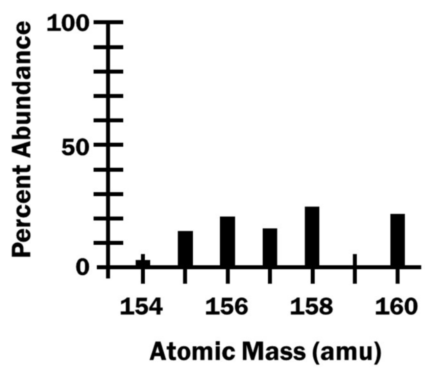
(b) Fill in the table below.

Mass Number	Protons	Neutrons
63		
65		

(c) Which isotope of this element, mass number = 63 or mass number = 65, is more abundant in nature? Justify your answer.

11. If an element has several naturally occurring isotopes, the calculation of the average atomic mass of the element can be a bit more complicated.

Mass Number	Abundance
154	2.18%
155	14.80%
156	20.47%
157	15.65%
158	24.84%
160	22.06%



(a) Based on the information above, estimate the average atomic mass of the element to the nearest whole number. Then use a calculator to determine the average atomic mass.

(b) What is the most likely identity of this element? ______________________________

1.3 Elemental Composition of Pure Substances

Essential knowledge statements from the AP Chemistry CED:

- Some pure substances are composed of individual molecules, while others consist of atoms or ions held together in fixed proportions as described by a formula unit.
- According to the law of definite proportions, the ratio of the masses of the constituent elements in any pure sample of that compound is always the same.
- The chemical formula that lists the lowest whole number ratio of atoms of the elements in a compound is the empirical formula.

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12. Calculate the percent composition by mass of each element in glucose ($C_6H_{12}O_6$).

13. Calculate the percent composition by mass of each element in erythrose ($C_4H_8O_4$).

14. What is the empirical formula of glucose? ____________________

    What is the empirical formula of erythrose? ___________________

Two different compounds with the same empirical formula have the same percent composition by mass.

15. A certain compound has the following percent composition by mass.

    43.64% P   56.36% O

    Determine the empirical formula of this compound.

If you are given mass data for a certain compound, the following procedure will help you to determine the empirical formula of the compound.

- Convert the mass of each element into moles.
- Divide each value of moles by the lowest number.
- At this point, you may already have whole numbers for the moles of each element. If not, then you may need to multiply each value by 2 or by 3 in order to get whole numbers.
- Use the whole number values of moles to write the empirical formula.

16. A certain compound has the following percent composition by mass.

52.14% C   13.13% H   34.73% O

Determine the empirical formula of this compound.

17. A pure sample of tin (Sn) with a mass of 6.18 g is burned in air until the tin is completely converted into tin oxide. The mass of the tin oxide is equal to 7.85 g. Determine the empirical formula of the tin oxide compound.

18. Compound X consists of the elements C, H, and N. A 15.00-g sample of compound X contains 9.81 g C, 1.37 g H, and 3.82 g N.

(a) Determine the empirical formula of compound X.

(b) It is determined that a 25.0-gram sample of compound X contains 9.11×10^{22} molecules. Calculate the molar mass of compound X, in units of g/mol.

18. (c) Based on your answers to parts (a) and (b), determine the molecular formula of compound X.

Another way to determine the empirical formula of a compound is to use data from a combustion experiment. If a compound that contains carbon and hydrogen is burned in the presence of excess oxygen gas, the carbon will be converted into CO_2 and the hydrogen will be converted into H_2O. If the compound contains other elements such as nitrogen or sulfur, other gases may be formed.

Mass of sample that is burned	5.00 g
Mass of CO_2 produced	10.99 g
Mass of H_2O produced	6.00 g

19. A sample of a compound that contains carbon, hydrogen, and oxygen is burned completely in O_2. Data from the combustion experiment is shown in the table above.

    (a) Determine the mass of carbon (C) present in 5.00 g of the compound.

    (b) Determine the mass of hydrogen (H) present in 5.00 g of the compound.

    (c) Determine the mass of oxygen (O) present in 5.00 g of the compound.

    (d) Determine the empirical formula of the compound.

Another type of situation that involves mass and mole ratios involves a substance known as a hydrate. A hydrate is a substance in which water molecules are included in the chemical formula. These substances are often ionic compounds in which water molecules are bonded to the ions in the crystal structure. A hydrate salt can be heated to remove the water through evaporation, forming an anhydrous salt. Two examples of anhydrous salts and hydrates are listed in the table below.

Anhydrous Salt	Hydrate Salt
copper(II) sulfate, $CuSO_4$	copper(II) sulfate pentahydrate, $CuSO_4{\bullet}5H_2O$
calcium chloride, $CaCl_2$	calcium chloride dihydrate, $CaCl_2{\bullet}2H_2O$

20. A sample of $CuSO_4{\bullet}5H_2O$ has a mass of 25.00 g.

    (a) Calculate the mass of $CuSO_4$ present in this 25.00-g sample.

    (b) Calculate the mass of H_2O present in this 25.00-g sample.

21. Calculate the percentage of H_2O by mass in $CaCl_2{\bullet}2H_2O$.

22. In a certain experiment, a sample of a hydrate of magnesium sulfate, $MgSO_4{\bullet}nH_2O$, is heated in order to remove all of the water from the sample. Experimental data is shown in the table below.

mass of empty container	25.356 g
mass of container and hydrate salt, before heating	28.418 g
mass of container and sample after 1st heating	26.931 g
mass of container and sample after 2nd heating	26.853 g
mass of container and sample after 3rd heating	26.852 g

    (a) Explain how the data indicates that all of the water has been removed from the hydrate salt in this experiment.

22. (b) Calculate the mass of the hydrate salt used in this experiment.

(c) Calculate the mass of water that was removed from the hydrate sample in this experiment.

(d) Determine the value of n in the formula $MgSO_4 \bullet nH_2O$.