The mass number of an isotope equals the total number of protons and neutrons in an atom which can be obtained by rounding the atomic mass in the periodic table to the nearest whole number. This is not the same as the atomic mass, which is an average of the masses of the isotopes of the particular element.

WRITING SYMBOLS

(*Chemistry* 6th ed. page 54 / 7th ed. pages 50–51)

Mass number → $^{23}_{11}Na$
Atomic Mass →

EXAMPLE: How many protons, neutrons and electrons are in an atom of sodium – 23?

SOLUTION: There are 11 protons since the atomic number is 11. There are also 11 electrons since atoms are neutral. There are 12 neutrons (mass number – atomic number = 23 – 11).

ISOTOPES

(*Chemistry* 6th ed. pages 81–85 / 7th ed. pages 78–81)

Isotopes are atoms of the same element with different numbers of neutrons and therefore different atomic masses.

CALCULATION OF AVERAGE ATOMIC MASS FROM ISOTOPIC DATA

(*Chemistry* 6th ed. pages 81–85 / 7th ed. pages 78–81)

The average atomic mass of an element can be calculated from the percent abundance and mass of each isotope for that element.

Avg Atomic Mass = Σ(Percent abundance of isotope × mass value)

EXAMPLE: Element "E" is present with the following mass values and natural abundances.

Isotope	Mass Value (amu)	Percent Abundance
10 E	10.01	19.78 %
11 E	11.01	80.22 %

What is the average atomic mass of the element, E? What is the element?

(0.1978) 10.01 + (0.8022) 11.01 = 10.812 amu.

The element is boron.

IONS

(*Chemistry* 6th ed. pages 57, 62, 63, 67 / 7th ed. pages 53, 57–58, 62)

An ion is an atom that has lost or gained electrons. A polyatomic ion is a group of atoms bonded together as a single unit that carries a charge.

EXAMPLE: Aluminum forms a cation, a positive ion, by losing three electrons:

$Al \rightarrow Al^{3+} + 3e^-$

Oxygen forms an anion, a negative ion, by gaining three electrons:

$O + 2e^- \rightarrow O^{2-}$

Ch. 3

STOICHIOMETRY

Stoichiometry includes calculations of average atomic mass, the mole, molar mass, percent composition, molarity and empirical formula. The quantities of materials consumed and produced in chemical reactions are also considered. The multiple-choice section of the AP Exam may contain questions that require computations without calculators. The free response portion of the exam may contain one multistep question on stoichiometry.

You should understand

- **Moles, mass, representative particles (atoms, molecules, formula units), molar mass, and Avogadro's number.**
- **Molarity; preparation of solutions.**
- **The percent composition of an element in a compound.**
- **Balanced chemical equations; for example, for a given mass of reactant, calculate the amount of product produced.**
- **Limiting reactants: calculate the amount of product formed when given the amounts of all of the reactants present.**
- **Reactions in solution: given the molarity and the volume of the reactants, calculate the amount of product produced or the amount of reactant required to react.**
- **The percent yield of a reaction.**

THE MOLE

(*Chemistry* 6th ed. pages 86–90 / 7th ed. pages 82–85)

The mole is defined in the table below. Calculations involving moles utilize the conversions in the table below.

1 mol of a monoatomic element = 6.02×10^{23} atoms
1 mol of a molecular compound or diatomic element = 6.02×10^{23} molecules
1 mol of an ionic compound = 6.02×10^{23} formula units

MOLAR MASS

(*Chemistry* 6th ed. pages 90–93 / 7th ed. pages 86–88)

Molar mass is the mass of 1 mole of an element or compound.

EXAMPLE: The principal ore in the production of aluminum cans has a molecular formula of $Al_2O_3 \cdot 2H_2O$. What is the mass in grams of 2.10×10^{24} formula units (f.u.) of $Al_2O_3 \cdot 2H_2O$?

$$2.10\times10^{24}\text{ formula units}\times\frac{1\text{ mol }Al_2O_3\cdot2H_2O}{6.02\times10^{23}\text{ f.u.}}\times\frac{138\text{ g}}{1\text{ mol }Al_2O_3\cdot2\,H_2O}=481\text{ g}$$

MOLARITY

(*Chemistry* 6th ed. pages 140–146 / 7th ed. pages 133–139)

The number of moles of solute in 1 liter of solution is a measure of its molarity or solution concentration.

EXAMPLE: Prepare 2.00 L of 0.250 M NaOH from solid NaOH.

$$2.00\text{ L}\times\frac{0.250\text{ mol NaOH}}{\text{L}}\times\frac{40.00\text{ g NaOH}}{1\text{ mol}}=\mathbf{20.0\text{ g NaOH}}$$

Place 20.0 g NaOH in a 2-L volumetric flask; add water to dissolve the NaOH, and fill to the mark with water, mixing several times along the way.

EXAMPLE: Prepare 2.00 L of 0.250 *M* NaOH from 1.00 *M* NaOH.

$$\boxed{M_1V_1 = M_2V_2}$$

$1.00\,M\ \ V_1 = 0.250\,M \times 2.00\text{ L}$ $\qquad$ $\mathbf{V_1 = 0.500\text{ L}}$

Add 500. mL of 1.00 *M* NaOH stock solution to a 2-L volumetric flask; add deionized water in several increments, mixing until the flask is filled to the mark on the neck of the flask.

PERCENT COMPOSITION

(*Chemistry* 6th ed. pages 93–96 / 7th ed. pages 89–91)

The mass percentages of elements in a compound can be obtained by comparing the mass of each element present in 1 mole of the compound to the total mass of 1 mole of compound.

$$\boxed{\frac{\text{Mass of element in 1 mol of compound}}{\text{Mass of 1 mol of compound}}\times100\%}$$

EXAMPLE: Calculate the percent of oxygen in $Mg(NO_3)_2$
= (16 × 6/148) × 100% = **65%**

DETERMINATION OF EMPIRICAL FORMULA BY COMBUSTION ANALYSIS

(*Chemistry* 6th ed. pages 96–98 / 7th ed. pages 91–93)

The empirical formula is the simplest whole number ratio of atoms in a compound.

(The following problem involves two steps. First, the percent composition of the compound is determined from combustion data. Second, the empirical formula of the compound is determined from the percent composition.)

EXAMPLE: A compound contains only carbon, hydrogen, and oxygen. Combustion of 10.68 mg of the compound yields 16.01 mg CO_2 and 4.37 mg of H_2O. What is the percent composition of the compound?

SOLUTION: Assume that all the carbon in the compound is converted to CO_2 and determine the mass of carbon present in the 10.68-mg sample.

$$.01601\text{ g }CO_2\times\frac{12.01\text{ g C}}{44.01\text{ g }CO_2}=4.369\text{ mg C}$$

The mass percent of C in this compound is

$$\frac{4.369\text{ mg C}}{10.68\text{ mg}}\times100\%=40.91\%\text{ C}$$

The same procedure can be used to find the mass percent of hydrogen in the unknown compound. We assume that all the hydrogen present in 10.68 mg of compound was converted to H_2O.

$$0.0437\text{ g }H_2O\times\frac{2.016\text{ g H}}{18.02\text{ g }H_2O}=0.489\text{ mg H}$$