CHAPTER 3 REVIEW

CHAPTER SUMMARY

3-1

- The *idea* of atoms has been around since the time of the ancient Greeks. In the nineteenth century, John Dalton proposed a *scientific theory* of atoms that can still be used to explain properties of many chemicals today.
- When elements react to form compounds, they combine in fixed proportions by mass.
- Matter and its mass cannot be created or destroyed in chemical reactions.
- The mass ratio of the elements that make up a given compound is always the same, regardless of how much of the compound there is or how it was formed.
- If two or more different compounds are composed of the same two elements, then the ratio of the masses of the second element combined with a certain mass of the first element can be expressed as a ratio of small whole numbers.

Vocabulary

law of conservation of mass (66) law of definite proportions (66) law of multiple proportions (66)

3-2

- Cathode-ray tubes supplied evidence of the existence of electrons, which are negatively charged subatomic particles that have relatively little mass.
- Rutherford found evidence for the existence of the atomic nucleus—a positively charged, very dense core within the atom—by bombarding metal foil with a beam of positively charged particles.
- Atomic nuclei are composed of protons, which have an electric charge of +1, and (in all but one case) neutrons, which have no electric charge.
- Isotopes of an element differ in the number of neutrons in their nuclei.
- Atomic nuclei have radii of about 0.001 pm (pm = picometers; 1 pm = 10^{-12} m), while atoms have radii of about 40–270 pm.

Vocabulary

atom (70) nuclear forces (74)

3-3

- The atomic number of an element is equal to the number of protons in the nucleus of an atom of that element.
- The mass number is equal to the total number of protons and neutrons in the nucleus of an atom of that element.
- The relative atomic mass unit (amu) is based on the carbon-12 atom and is a convenient unit for measuring the mass of atoms. It equals $1.660\ 540 \times 10^{-24}$ g.
- The average atomic mass of an element is found by calculating the weighted average of the atomic masses of the naturally occurring isotopes of the element.
- Avogadro's number is equal to approximately $6.022\ 137 \times 10^{23}$. It is equal to the number of atoms in exactly 12 g of carbon-12. A sample that contains a number of particles equal to Avogadro's number contains a mole of those particles.
- The molar mass of an element is the mass of one mole of atoms of that element.

Vocabulary

atomic mass unit (78)
atomic number (75)
average atomic mass (79)
Avogadro's number (81)
isotopes (76)
mass number (76)
molar mass (81)
mole (81)
nuclide (77)