Math Tutor | Conversion Factors

Most calculations in chemistry require that all measurements of the same quantity (mass, length, volume, temperature, and so on) be expressed in the same unit. To change the units of a quantity, you can multiply the quantity by a conversion factor. With SI units, such conversions are easy because units of the same quantity are related by powers of 10, 100, 1000, or 1 million. Suppose you want to convert a given amount in milliliters to liters. You can use the relationship 1 L = 1000 mL. From this relationship, you can derive the following conversion factors.

$$\frac{1000\ \text{mL}}{1\ \text{L}} \text{ and } \frac{1\ \text{L}}{1000\ \text{mL}}$$

The correct strategy is to multiply the given amount (in mL) by the conversion factor that allows milliliter units to cancel out and liter units to remain. Using the second conversion factor will give you the units you want.

These conversion factors are based on an exact definition (1000 mL = 1 L exactly), so significant figures do not apply to these factors. The number of significant figures in a converted measurement depends on the certainty of the measurement you start with.

Sample Problem

A sample of aluminum has a mass of 0.087 g. What is the sample's mass in milligrams?

Based on SI prefixes, you know that 1 g = 1000 mg. Therefore, the possible conversion factors are

$$\frac{1000\ \text{mg}}{1\ \text{g}} \text{ and } \frac{1\ \text{g}}{1000\ \text{mg}}$$

The first conversion factor cancels grams, leaving milligrams.

$$0.087\ \cancel{\text{g}} \times \frac{1000\ \text{mg}}{1\ \cancel{\text{g}}} = 87\ \text{mg}$$

Notice that the values 0.087 g and 87 mg each have two significant figures.

A sample of a mineral has 4.08×10^{-5} mol of vanadium per kilogram of mass. How many micromoles of vanadium per kilogram does the mineral contain?

The prefix *micro-* specifies $\frac{1}{1\,000\,000}$, or 1×10^{-6}, of the base unit. So, 1 μmol = 1×10^{-6} mol. The possible conversion factors are

$$\frac{1\ \mu\text{mol}}{1 \times 10^{-6}\ \text{mol}} \text{ and } \frac{1 \times 10^{-6}\ \text{mol}}{1\ \mu\text{mol}}$$

The first conversion factor will allow moles to cancel and micromoles to remain.

$$4.08 \times 10^{-5}\ \cancel{\text{mol}} \times \frac{1\ \mu\text{mol}}{1 \times 10^{-6}\ \cancel{\text{mol}}} = 40.8\ \mu\text{mol}$$

Notice that the values 4.08×10^{-5} mol and 40.8 μmol each have three significant figures.

Practice

1. Express each of the following measurements in the units indicated.
   - **a.** 2250 mg in grams
   - **b.** 59.3 kL in liters
2. Use scientific notation to express each of the following measurements in the units indicated.
   - **a.** 0.000 072 g in micrograms
   - **b.** 3.98×10^{6} m in kilometers

CHAPTER 3 Summary

BIG IDEA Atoms are the smallest particles of elements. The atoms of one element differ from the atoms of another by the number of protons they contain.

SECTION 1 The Atom: From Philosophical Idea to Scientific Theory

- 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 most chemicals today.
- Matter and its mass cannot be created or destroyed in chemical reactions.
- The mass ratios of the elements that make up a given compound are 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.

KEY TERMS

law of conservation of mass
law of definite proportions
law of multiple proportions

SECTION 2 The Structure of the Atom

- 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 by bombarding gold 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.
- Atomic nuclei have radii of about 0.001 pm (pm = picometers; 1 pm = 10^{-12} m), and atoms have radii of about 40–270 pm.

KEY TERMS

atom
nuclear forces

SECTION 3 Counting Atoms

- The atomic number of an element is equal to the number of protons of an atom of that element.
- The mass number is equal to the total number of protons and neutrons that make up the nucleus of an atom of that element.
- The unified atomic mass unit (u) 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 6.022×10^{23}. A sample that contains a number of particles equal to Avogadro's number contains a mole of those particles.

KEY TERMS

atomic number
isotope
mass number
nuclide
unified atomic mass unit
average atomic mass
mole
Avogadro's number
molar mass

CHAPTER 3 Review

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SECTION 1

The Atom: From Philosophical Idea to Scientific Theory

REVIEWING MAIN IDEAS

1. Explain each of the following in terms of Dalton's atomic theory:
   a. the law of conservation of mass
   b. the law of definite proportions
   c. the law of multiple proportions

2. According to the law of conservation of mass, if element A has an atomic mass of 2 mass units and element B has an atomic mass of 3 mass units, what mass would be expected for compound AB? for compound A_2B_3?

SECTION 2

The Structure of the Atom

REVIEWING MAIN IDEAS

3. a. What is an atom?
   b. What two regions make up all atoms?

4. Describe at least four properties of electrons that were determined based on the experiments of Thomson and Millikan.

5. Summarize Rutherford's model of the atom, and explain how he developed this model based on the results of his famous gold-foil experiment.

6. What number uniquely identifies an element?

SECTION 3

Counting Atoms

REVIEWING MAIN IDEAS

7. a. What are isotopes?
   b. How are the isotopes of a particular element alike?
   c. How are they different?

8. Copy and complete the following table concerning the three isotopes of silicon, Si. (Hint: See Sample Problem A.)

Isotope	Number of protons	Number of electrons	Number of neutrons
Si-28			
Si-29			
Si-30			

9. a. What is the atomic number of an element?
   b. What is the mass number of an isotope?
   c. In the nuclear symbol for deuterium, $^{2}_{1}H$, identify the atomic number and the mass number.

10. What is a nuclide?

11. Use the periodic table and the information that follows to write the hyphen notation for each isotope described.
   a. atomic number = 2, mass number = 4
   b. atomic number = 8, mass number = 16
   c. atomic number = 19, mass number = 39

12. a. What nuclide is used as the standard in the relative scale for atomic masses?
   b. What is its assigned atomic mass?

13. What is the atomic mass of an atom if its mass is approximately equal to the following?
   a. $\frac{1}{3}$ that of carbon-12
   b. 4.5 times as much as carbon-12

14. a. What is the definition of a *mole*?
   b. What is the abbreviation for *mole*?
   c. How many particles are in one mole?
   d. What name is given to the number of particles in a mole?

15. a. What is the molar mass of an element?
   b. To two decimal places, write the molar masses of carbon, neon, iron, and uranium.

16. Suppose you have a sample of an element.
   a. How is the mass in grams of the element converted to amount in moles?
   b. How is the mass in grams of the element converted to number of atoms?

PRACTICE PROBLEMS

17. What is the mass in grams of each of the following? (Hint: See Sample Problems B and E.)
- **a.** 1.00 mol Li
- **b.** 1.00 mol Al
- **c.** 1.00 molar mass Ca
- **d.** 1.00 molar mass Fe
- **e.** 6.022×10^{23} atoms C
- **f.** 6.022×10^{23} atoms Ag

18. How many moles of atoms are there in each of the following? (Hint: See Sample Problems C and D.)
- **a.** 6.022×10^{23} atoms Ne
- **b.** 3.011×10^{23} atoms Mg
- **c.** 3.25×10^{5} g Pb
- **d.** 4.50×10^{-12} g O

19. Three isotopes of argon occur in nature—$^{36}_{18}\text{Ar}$, $^{38}_{18}\text{Ar}$, and $^{40}_{18}\text{Ar}$. Calculate the average atomic mass of argon to two decimal places, given the following relative atomic masses and abundances of each of the isotopes: argon-36 (35.97 u; 0.337%), argon-38 (37.96 u; 0.063%), and argon-40 (39.96 u; 99.600%).

20. Naturally occurring boron is 80.20% boron-11 (atomic mass = 11.01 u) and 19.80% of some other isotopic form of boron. What must the atomic mass of this second isotope be in order to account for the 10.81 u average atomic mass of boron? (Write the answer to two decimal places.)

21. How many atoms are there in each of the following?
- **a.** 1.50 mol Na
- **b.** 6.755 mol Pb
- **c.** 7.02 g Si

22. What is the mass in grams of each of the following?
- **a.** 3.011×10^{23} atoms F
- **b.** 1.50×10^{23} atoms Mg
- **c.** 4.50×10^{12} atoms Cl
- **d.** 8.42×10^{18} atoms Br
- **e.** 25 atoms W
- **f.** 1 atom Au

23. Determine the number of atoms in each of the following:
- **a.** 5.40 g B
- **b.** 0.250 mol S
- **c.** 0.0384 mol K
- **d.** 0.025 50 g Pt
- **e.** 1.00×10^{-10} g Au

Mixed Review

REVIEWING MAIN IDEAS

24. Determine the mass in grams of each of the following:
- **a.** 3.00 mol Al
- **b.** 2.56×10^{24} atoms Li
- **c.** 1.38 mol N
- **d.** 4.86×10^{24} atoms Au
- **e.** 6.50 mol Cu
- **f.** 2.57×10^{8} mol S
- **g.** 1.05×10^{18} atoms Hg

25. Copy and complete the following table concerning the properties of subatomic particles.

Particle	Symbol	Mass number	Actual mass	Relative charge
Electron				
Proton				
Neutron				

26.
- **a.** How is a unified atomic mass unit (u) related to the mass of one carbon-12 atom?
- **b.** What is the relative atomic mass of an atom?

27.
- **a.** What is the nucleus of an atom?
- **b.** Who is credited with the discovery of the atomic nucleus?
- **c.** Identify the two kinds of particles that make up the nucleus.

28. How many moles of atoms are there in each of the following?
- **a.** 40.1 g Ca
- **b.** 11.5 g Na
- **c.** 5.87 g Ni
- **d.** 150 g S
- **e.** 2.65 g Fe
- **f.** 0.007 50 g Ag
- **g.** 2.25×10^{25} atoms Zn
- **h.** 50 atoms Ba

29. State the law of multiple proportions, and give an example of two compounds that illustrate the law.

30. What is the approximate atomic mass of an atom if its mass is
- **a.** 12 times that of carbon-12?
- **b.** $\frac{1}{2}$ that of carbon-12?

31. What is an electron?

CRITICAL THINKING

32. Organizing Ideas Using two chemical compounds as an example, describe the difference between the law of definite proportions and the law of multiple proportions.

33. Constructing Models As described in Section 2, the structure of the atom was determined from observations made in painstaking experimental research. Suppose a series of experiments revealed that when an electric current is passed through gas at low pressure, the surface of the cathode-ray tube opposite the anode glows. In addition, a paddle wheel placed in the tube rolls from the anode toward the cathode when the current is on.

a. In which direction do particles pass through the gas?

b. What charge do the particles possess?

34. Analyzing Data Osmium is the element with the greatest density, 22.58 g/cm^3. How does the density of osmium compare to the density of a typical nucleus of 2×10^8 metric tons/cm^3? (1 metric ton = 1000 kg)

USING THE HANDBOOK

35. Group 14 of the *Elements Handbook* (Appendix A) describes the reactions that produce CO and CO_2. Review this section to answer the following:

a. When a fuel burns, what determines whether CO or CO_2 will be produced?

b. What happens in the body if hemoglobin picks up CO?

c. Why is CO poisoning most likely to occur in homes that are well sealed during cold winter months?

RESEARCH AND WRITING

36. Prepare a report on the series of experiments conducted by Sir James Chadwick that led to the discovery of the neutron. In your report, evaluate the impact of this research on scientific thought about the structure of the atom.

37. Write a report on the contributions of Amedeo Avogadro that led to the determination of the value of Avogadro's number.

38. Trace the development of the electron microscope, and cite some of its many uses.

39. The study of atomic structure and the nucleus produced a new field of medicine called *nuclear medicine*. Describe the use of radioactive tracers to detect and treat diseases.

ALTERNATIVE ASSESSMENT

40. Observe a cathode-ray tube in operation, and write a description of your observations.

41. Performance Assessment Using colored clay, build a model of the nucleus of each of carbon's three naturally occurring isotopes: carbon-12, carbon-13, and carbon-14. Specify the number of electrons that would surround each nucleus.