Measurement and Problem Solving   2

Chapter Overview

Chapter 2 introduces the student to a cornerstone of the chemical sciences, the manipulation of numbers and their associated units. These concepts are very important for the rest of the course, and in order to be successful in this course, students must understand them well. Simple and complex unit conversions as well as problem-solving strategies will be covered and explained in detail.

Lecture Outline

2.1 Measuring Global Temperatures
    A. Units are important
    B. How many digits do I report?
2.2 Scientific Notation: Writing Large and Small Numbers
    Learning Objective: Express very large and very small numbers using scientific notation.
    A. Shorthand notation for numbers
    B. Two main pieces: decimal and power-of-10 exponent
    C. Measured value does not change, just how you report it
2.3 Scientific Figures: Writing Numbers to Reflect Precision
    Learning Objective: Report measured quantities to the right number of digits.
    A. How many digits can I report? How many should I report?
    B. Certain digits and estimated digits
    C. Counting significant figures
        1. All nonzero digits are significant
        2. Interior zeros are significant
        3. Trailing zeros after a decimal point are significant
        4. Trailing zeros before a decimal point are significant
        5. Leading zeros are not significant
        6. Zeros at the end of a number, but to the left of a decimal point, are ambiguous
    D. Exact numbers
2.4 Significant Figures in Calculations
    Learning Objective: Round numbers to the correct number of significant figures.
    Learning Objective: Determine the correct number of significant figures in the results of multiplication and division calculations.
    Learning Objective: Determine the correct number of significant figures in the results of addition and subtraction calculations.

Learning Objective: Determine the correct number of significant figures in the results of calculations involving both addition/subtraction and multiplication/division.

A. Multiplication and Division
     1. Result carries as many significant digits as the factor with the fewest significant digits.

B. Rounding
     1. If leftmost dropped digit is 4 or less, round down
     2. If leftmost dropped digit is 5 or higher, round up

C. Addition and Subtraction
     1. Result carries as many decimal places as the quantity with the fewest decimal places

D. Calculations Involving Both Multiplication/Division and Addition/Subtraction
     1. Do steps in parentheses first
     2. Determine the number of significant figures in intermediate answer
     3. Do remaining steps

2.5 The Basic Units of Measurement

Learning Objective: Recognize and work with the SI base units of measurement, prefix multipliers, and derived units.

A. English, metric, SI

B. SI Units
     1. Length – m
     2. Mass – kg
     3. Time – s

C. Prefix Multipliers
     1. milli (m) 0.001
     2. centi (c) 0.01
     3. kilo (k) 1000
     4. Mega (M) 1,000,000

D. Derived Units
     1. Area – cm^2
     2. Volume – cm^3 or L

2.6 Problem Solving and Unit Conversions

Learning Objective: Convert between units.

A. Units are important, most numbers have one

B. Include units in all calculations

C. Conversion factors change one unit to another, the value is unchanged

2.7 Solving Multistep Conversion Problems

Learning Objective: Convert between units.

A. Understand where you are going first

B. Not all calculations can be done in one step

2.8 Units Raised to a Power

Learning Objective: Convert units raised to a power.

A. 1 inch = 2.54 cm so 1 $inch^3$ = $(2.54)^3$ cm^3 = 16.4 cm^3

2.9 Density

Learning Objective: Calculate the density of a substance.

Learning Objective: Use density as a conversion factor.

A. Mass per unit volume
B. Derived unit
C. Can be used as a conversion factor between mass and volume

2.10 Numerical Problem-Solving Strategies and the Solution Map
A. Come up with a plan before you use your calculator
B. Use the units to guide your plan

Chemical Principle Teaching Ideas

Uncertainty

Students generally have a hard time understanding this concept. One method is to refer to everyday objects that they recognize. For example, you can talk about a coffee cup containing about 200 mL of coffee. You then ask the students what the new volume would be if you were to add a drop of water with a volume of 0.05 mL.

Units

Units are very important, and should always be used. Consider giving the students a measured value in many different units and having them guess what the unit is. Report the volume of your mug in barrels. What is the volume of the room measured in teaspoons?

Density

Most students understand the concept of density, or how much stuff is packed into a particular volume. What they have a harder time recognizing is the fact that it is a conversion factor between mass and volume. This is the easiest example that is discussed and should be emphasized as this concept is used frequently throughout the course.

Skill Builder Solutions

2.1.   Assuming all the trailing zeros are not significant, the decimal moves over 13 spaces to give $\$1.6342 \times 10^{13}$.

2.2.   All the leading zeros are not significant, so we move the decimal over 5 places to give 3.8×10^{-5}.

2.3.   Each of the markings on the thermometer represents 1 degree Fahrenheit. We can therefore estimate one digit past the decimal place for a temperature of 103.4 degrees Fahrenheit.

2.4.   a. 4
        b. 3, as leading zeros do not count, but trailing zeros after the decimal do
        c. 2
        d. Unlimited significant figures
        e. 3
        f. Ambiguous, since you do not know if the last 2 zeros are significant

2.5. a. $\dfrac{1.10 \times 0.512 \times 1.301 \times 0.005}{3.4} = 0.001$. There is only one significant digit in the final

answer as the 0.005 has only one significant digit in the numerator.

b. $\dfrac{4.562 \times 3.99870}{89.5} = 0.204$. The number 89.5 has the fewest number of significant

digits, 3, so that is how many quoted in the final answer.

2.6. a. $2.18 + 5.621 + 1.5870 - 1.8 = 7.6$. Only one digit past the decimal place is quoted
   because the least accurately known number (1.8) has one digit past the decimal.
   b. $7.876 - 0.56 + 123.792 = 131.11$. Two digits past the decimal are quoted, because 0.56
   has two past the decimal and is the number with the fewest digits past the decimal.

2.7. a. $3.897 \times (782.3 - 451.88) = 3.897 \times 330.\underline{4}2 = 1288$. Four digits are quoted because the
   number in the second (multiplication) step with the fewest significant digits has four of
   them.
   b. $\dfrac{4.58}{1.239} - 0.578 = 3.70 - 0.578 = 3.12$. Two digits past the decimal are quoted because

the first part of the subtraction (3.70) has two digits past the decimal place.

2.8. $56.0 \ \cancel{cm} \times \dfrac{1 \text{ inch}}{2.54 \ \cancel{cm}} = 22.0 \text{ inch}$

2.9. $5{,}678 \ \cancel{m} \times \dfrac{1 \text{ km}}{1000 \ \cancel{m}} = 5.678 \text{ km}$

2.10. $1.2 \ \cancel{cu} \times \dfrac{1 \ \cancel{qt}}{4 \ \cancel{cu}} \times \dfrac{1 \text{ L}}{1.057 \ \cancel{qt}} = 0.28 \text{ L}$

2.11. $15.0 \ \cancel{km} \times \dfrac{0.6214 \ \cancel{mi}}{1 \ \cancel{km}} \times \dfrac{5280 \ \cancel{ft}}{1 \ \cancel{mi}} \times \dfrac{1 \text{ lap}}{1056 \ \cancel{ft}} = 46.6 \text{ laps}$

Plus. $5.72 \ \cancel{naut\,mi} \times \dfrac{1.151 \ \cancel{mi}}{1 \ \cancel{naut\,mi}} \times \dfrac{1 \ \cancel{km}}{0.6214 \ \cancel{mi}} \times \dfrac{1000 \text{ m}}{1 \ \cancel{km}} = 1.06 \times 10^4 \text{ m}$

2.12. $289.7 \ \cancel{in^3} \times \dfrac{(2.54)^3 \text{ cm}^3}{1 \ \cancel{in^3}} = 4747 \text{ cm}^3$

2.13. $3.25 \ \cancel{yd^3} \times \dfrac{(36)^3 \text{ inch}^3}{1 \ \cancel{yd^3}} = 1.52 \times 10^5 \text{ inch}^3$

2.14. $\dfrac{9.67 \text{ g}}{0.452 \text{ cm}^3} = 21.4 \text{ g/cm}^3$; Therefore the ring is genuine platinum.

2.15. $35 \text{ mg} \times \dfrac{1 \text{ g}}{1000 \text{ mg}} \times \dfrac{1 \text{ cm}^3}{0.788 \text{ g}} = 4.4 \times 10^{-2} \text{ cm}^3$

Plus. $246 \text{ cm}^3 \times \dfrac{7.93 \text{ g}}{1 \text{ cm}^3} \times \dfrac{1 \text{ kg}}{1000 \text{ g}} = 1.95 \text{ kg}$

2.16. $0.82 \text{ L} \times \dfrac{1000 \text{ mL}}{1 \text{ L}} \times \dfrac{19.3 \text{ g}}{1 \text{ mL}} \times \dfrac{1 \text{ kg}}{1000 \text{ g}} = 16 \text{ kg}$

2.17. $\dfrac{23.2 \text{ mg} \times \dfrac{1 \text{ g}}{1000 \text{ mg}}}{1.20 \text{ mm}^3 \times \dfrac{1 \text{ cm}^3}{(10)^3 \text{ mm}^3}} = \dfrac{2.32 \times 10^{-2} \text{ g}}{1.20 \times 10^{-3} \text{ cm}^3} = 19.3 \text{ g/cm}^3$

Yes, it is consistent with the density of gold.

Suggested Demonstrations

Density and Miscibility of Liquids, *Chemical Demonstrations* 3:233, Shakhashiri, B.Z. University of Wisconsin Press, 1989.

Guided Inquiry Ideas

Below are a few example questions that students answer in the guided inquiry activities provided in the Guided Activity Workbook.

How many significant figures are there in the number 0.0051? Underline it/them.

How many significant figures are there in the number 5.00? Underline it/them.

In a complete sentence or two describe when you know a "trailing zero" is significant.

In a complete sentence, describe the significance of "leading zeros".

Which of the following is a correct conversion factor from cm^3 to in^3? Circle all that apply.

$\left(\dfrac{1 \text{ in}^3}{2.54 \text{ cm}^3} \right)$ $\qquad$ $\left(\dfrac{1 \text{ in}}{2.54 \text{ cm}} \right)^3$ $\qquad$ $\left(\dfrac{1 \text{ in}^3}{16.4 \text{ cm}^3} \right)$