CHEMICAL FOUNDATIONS

Ch. 1

In this chapter, you will perform the basic calculations used in chemistry. Chemists use equipment with varying degrees of precision. The measurements recorded in an experiment must reflect the precision of the equipment used. The results of the calculations in an experiment also must reflect the precision of the equipment.

Many of the problems in chemistry use dimensional analysis to convert from one unit to another or to solve stoichiometry problems. Other basic calculations that will be discussed are density and temperature conversions.

You should be able to

- Identify the number of significant figures in a given measurement.
- Perform calculations involving significant figures.
  - Memorize the rules for counting and performing operations with significant figures.
- Differentiate between measurements which are accurate and precise.
- Determine the density of solids and liquids and calculate volumes or masses using the given density.
- Convert between units of temperature: degrees Celsius and Fahrenheit, and Kelvin.
- Identify the characteristics of the states of matter: solids, liquids, and gases.
- Identify substances such as elements, compounds, or mixtures.
- Identify methods of separation of mixtures.
- Identify changes as being physical or chemical.

AP tips

Significant figures are important in calculations. One point out of the nine possible in the free-response sections will be deducted for errors in significant figures that are off by ± 1 significant figure.

Units are the key to problem solving. Use conversion factors and cancel out units or problems involving conversions or stochiometry. Every measured number or number calculated from measurements has a unit except for equilibrium constants, which by convention are reported without units.

UNCERTAINTY IN MEASUREMENT

(*Chemistry* 6th ed. pages 11–14)

SIGNIFICANT FIGURES

The significant figures of a measurement are all of the certain digits in a measurement and the first uncertain digit (estimated number). Students should be able to read measurements to the proper number of significant figures.

EXAMPLE: Figure 1.9 on page 11 in the Zumdahl 6th edition text and page 10 in the 7th edition shows the measurement of a volume of liquid using a buret. The *certain digits* in the measurement are the three numbers 20.1. The digit to the right of the one must be estimated by interpolating between the 0.1 mL marks. The measurement with *uncertainty* can be reported as 20.15 mL

PRECISION AND ACCURACY

Accuracy refers to the agreement of a particular value with a true value.

Precision refers to the degree of agreement among several measurements of the same quantity. The degree of precision refers to the number of digits that a measuring device permits one to measure. In a measuring device, all except the last digit, which is estimated, are certain. For example, a balance which measures to the nearest 0.0001 g is more precise than one that measures to the nearest 0.01 g.

$$\text{Percent Error} = \frac{\text{Experimental Value} - \text{Actual Value}}{\text{Actual Value}} \times 100\%$$

EXAMPLE: In an experiment, the density of aluminum is to be determined. Two students perform the experiment three times and obtain the following results.

Trial	Student A	Student B
1	2.45 g/mL	2.69 g/mL
2	2.43 g/mL	2.70 g/mL
3	2.44 g/mL	2.71 g/mL

Describe the accuracy and precision of each student's results. For Student A, calculate the mean and the percent error if the actual value is 2.70 g / mL.

SOLUTION: The values for Student A are precise, but inaccurate. The values for student B are both precise and accurate.

Random error (indeterminate error) means that a measurement has an equal probability of being high or low.

Systematic error (determinate error) occurs in the same direction each time; it is either always high or always low.

EXAMPLE: A balance could have a defect causing it to give a result that is consistently 1.000 g too high.

SIGNIFICANT FIGURES AND CALCULATIONS

It is important to know the uncertainty in the final result in an experiment The final reported result cannot have more certainty than the least accurate measurement. The number of significant figures in a single value will be determined. Memorize the rules below and use them to answer the examples that follow.

Rules for Counting Significant Figures

1. Nonzero integers always count as significant figures.
2. Zeros: There are three classes of zeros.
   a. Leading zeros precede all the nonzero digits and do not count as significant figures. Example: 0.0025 has 2 significant figures.
   b. Captive zeros are zeros between nonzero numbers. These always count as significant figures. Example: 1.008 has 4 significant figures.
   c. Trailing zeros are zeros at the right end of the number.
   Trailing zeros are only significant if the number contains a decimal point. Example: 1.00×10^2 has three significant figures.
   Trailing zeros are not significant if the number does not contain a decimal point. Example: 100 has one significant figure.
3. Exact numbers, which can arise from counting or definitions such as 1 in = 2.54 cm, never limit the number of significant figures in a calculation.

EXAMPLE: How many significant figures are in each of the following?

100 L

SOLUTION: There is 1 significant figure. Trailing zeros do not count. Only the "1" is significant.

0.001010 L

SOLUTION: There are 4 significant figures. Leading zeros do not count. The numbers "1010" are significant because one zero is captive and the other zero is trailing, but a decimal is present.

Rules for Significant Figures in Calculations

1. For multiplication and division, the number of significant figures in the result is the same as the number with the least number of significant figures in the calculation.
2. For addition and subtraction, the result has the same number of decimal places as the number with the least number of decimal places in the calculation.
3. Rules for rounding:
   In a series of calculations, carry the extra digits to the final result, then round.
   If the digit to be removed
   - is less than 5, the preceding digit stays the same. For example, 2.44 rounds to 2.4.
   - is greater than or equal to 5, the preceding digit is increased by 1. For example, 2.45 rounds to 2.5.

It is important to calculate the results of mathematical expressions to the proper number of significant figures. Memorize the rules above and apply them to the examples that follow.

EXAMPLE: Perform the following calculations to the correct number of significant figures.

1) 16.8 g + 3.2557 g

SOLUTION: The calculator answer is 20.0557. The correct answer is 20.1g. The answer should have one decimal place.

2) 27 g / 4.148 mL

SOLUTION: The calculator answer is 6.509161041. The correct answer is 6.5 g/ mL. The answer should have two significant figures.

DIMENSIONAL ANALYSIS

(*Chemistry* 6th ed. pages 18–21)

Dimensional analysis is used to convert from one unit to another. It is the single most valuable mathematical technique that you will use in general chemistry. The method involves conversion factors to cancel units until you have the proper unit in the proper place. When you are setting up problems using dimensional analysis, you should be more concerned with units than numbers.

EXAMPLE: The density of mercury is 13.6 g/cm³. How many pounds would 1.00 liter of mercury weigh?

SOLUTION:

$$1.00\text{L} \times \frac{1000\ \text{cm}^3}{1\text{L}} \times \frac{13.6\ \text{g}}{1\ \text{cm}^3} \times \frac{1\ \text{pound}}{453.6\ \text{g}} = 30.0\ \text{pounds}$$