Technical Assistance Document

Use of Significant Digits in Chemistry

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Introduction

The 2010 *Chemistry Standards of Learning* and Curriculum Framework introduce the use of significant digits to students as a way of addressing uncertainty in measurements. The following chemistry standard and essential knowledge and skills from the Chemistry Curriculum Framework address the treatment of significant digits:

CH.1 The student will investigate and understand that experiments in which variables are measured, analyzed, and evaluated produce observations and verifiable data. Key concepts include

g) mathematical manipulations including SI units, scientific notation, linear equations, graphing, ratio and proportion, significant digits, and dimensional analysis.

Essential Knowledge and Skills

In order to meet this standard, it is expected that students will

- read measurements and record data, reporting the significant digits of the measuring equipment.

This document describes basic rules of significant digits and examples of significant digits in common calculations beyond the detail provided in the Chemistry Curriculum Framework. Students are expected to use significant digits in reporting the results of their calculations as much as is practical, including their use on Chemistry SOL assessments.

Background

All measurements* have some degree of uncertainty that come from a variety of sources. A statement of a measured value should include an estimate of the level of confidence associated with that value. Properly reporting an experimental result along with its uncertainty allows other people to make judgments about the quality of the experiment, and it facilitates meaningful comparisons with other similar values or theoretical predictions.

When making a measurement, it is generally assumed that some exact or true value exists based on how the measurement is defined. While the true value may never be known exactly, one attempts to find this ideal quantity to the best of his or her ability with the instruments available. When measuring by different methods, or even when making multiple measurements using the same method, slightly different results can be obtained. So how does one report the best estimate of this elusive true value? Further, how does one account for systematic error produced in multistep calculations involving many measurements?

Using significant digits is one way of tracking uncertainty and/or error. Every measurement is uncertain and has an error range, and those ranges are transmitted throughout calculations in which those measurements are used. Using significant digits provides an error range associated with every calculated value. Use of significant digits can help chemistry students compute as exactly as possible while being truthful about the limitations of their data.

*Percentages are also considered measurements.

Which digits are significant?

	Example	Number of Significant Digits
Nonzero numbers are always significant	12	2
Any zero between (captive) significant digits is significant	1043	4
All final zeros to the right of the decimal are significant	1.00	3
Placeholder zeros are not significant. To remove placeholder zeros, rewrite the number in scientific notation.	0.072 7.2×10^{-2}	2 2

	Example	Number of Significant Digits
Counted numbers and defined numbers (such as metric conversions) have an infinite number of significant zeroes. a. Counted numbers b. Defined numbers	a. 10 experiments, 3 apples, or 8 molecules b. 1 kilogram = 1,000 grams	a. an infinite number of significant digits b. an infinite number of significant digits

Using Significant Digits in Calculations

Operation	Rule	Example
Multiplication and Division	Round off the calculated result to the same number of significant digits as the measurement having the fewest significant digits.	11.079 g = 0.872362204 g/mL 12.7 mL *Answer with correct significant digits* **0.872 g/mL**
		6.6 (2 significant digits) x 7328.7 (5 significant digits) 48369.42 *Answer with correct significant digits* **4.8 x 10^4**
Addition and Subtraction	Round off the calculated result to the same number of decimal places as the measurement with the fewest decimal places. If there is not a decimal point, round the result back to the digit that is in the same position as the leftmost uncertain digit in the quantities being added or subtracted.	164 mL (3 significant digits) 18.16 mL (4 significant digits) + 39.7 mL (3 significant digits) 221.86 mL *Answer with correct significant digits* **222 mL**
		170 mL (2 significant digits) + 3.5 mL (2 significant digits) - 28 mL (2 significant digits) 145.5 mL *Answer with correct significant digits* **150 mL**
Mathematically exact numbers (Counting numbers, Conversion Factors, Definitions)	½ and 6.02 x 10^{23} are infinitely precise so they do not influence the precision of any computation. An inch is defined as exactly 2.54 cm so using the “defined” conversion factor in a calculation does not affect the number of significant digits.	10 marbles together have a mass of 265.7 g. What is the mass of each marble? 265.7 g/10=26.57 g *Answer with correct significant digits* **26.57 g** A mole of HCl = 36.461g HCl is also an exact number. 1.00794 g = H *(6 significant digits) + 35.453 g = Cl *(5 significant digits) 36.46094 g *Answer with significant digits* **36.461 g = HCl**

***** Based on the Periodic Table of Elements used for the Chemistry Standards of Learning Test