EXPERIMENT #1

WEIGHING AND VOLUMETRIC TECHNIQUES - ACCURACY AND PRECISION

Objectives

A. To become familiar with the proper techniques for using the analytic balance, burette, pipette and graduated cylinder.

B. To be able to distinguish between accuracy and precision in experimental results, and to determine the propagation of experimental uncertainty in calculations.

Introduction *(Zumdahl and Zumdahl, pp. 8-19 and pp. A10-A13)*

Uncertainty and Precision vs. Accuracy

You will encounter two types of numbers in scientific research. **Exact numbers** have specifically defined values, hence their values are known exactly. Examples of exact numbers include: exactly 12 eggs in a dozen; exactly 1000 mg in a gram; exactly 12 inches in one foot. We are certain that these numbers will always be the same, regardless of the conditions. In contrast, those numbers obtained through, or derived from, experimentation are sensitive towards the inherent limitations in the equipment or instrumentation used (as well as the individual(s) who collected the data!). In other words, *experimentally measured or derived quantities always have some degree of uncertainty*, and these quantities are referred to as **inexact numbers.** Thus, all quantities measured using any measuring device must be reported along with its uncertainty. Generally, the last digit in a measured number has some degree of uncertainty associated with it (*i.e.*, it is inexact). For example, the analytical balances used in this lab can be read to 0.0001 g, so all masses must be reported as, *e.g.*, 1.2345 ± 0.0001 g. Similarly, the 50 mL burettes can be read to 0.01 mL, so all burette readings must be reported as, *e.g.*, 12.34 ± 0.01 mL. In doing so, you are indicating the uncertainty of the last digit in the measured quantity ("5" for the former, "4" for the latter). The terms "precision" and "accuracy" are often used within the context of the uncertainties of measured values. Thus, **precision** is a measure of how closely individual measurements agree with one another, while **accuracy** refers to the degree of agreement between measured quantities and a "true" or "accepted" quantity (we'll discuss this more later).

Significant Figures

It is common sense that a calculation using measured quantities is governed by the least accurate measurement used in the calculation. That is, this particular measurement can be regarded as a limiting measurement. For this reason, it is important to know the number of **significant figures or digits** of a measured quantity (an inexact number). Note that exact numbers are regarded as having an infinite number of significant figures. In fact, when using exact numbers in calculations, we ignore their number of significant figures. In general, *the greater the number of significant figures, the greater is the certainty implied for the measurement.* Zeros are special digits because they may or may not be significant depending on

how they appear in the number. More specifically, *zeros may be used as part of the measured quantity, or they may be used to locate the position of a decimal point.* In the former instance, the zero is significant, while in the latter, not significant. Consider the following rules:

1. Zeros between non-zero digits are always significant, *e.g.*, 401 g and 4.01 g (both have 3 significant figures)

2. Zeros at the beginning of a number are never significant, and just provide the position of the decimal point, *e.g.*, 0.0012 g (2 significant figures) and 0.0134 mL (3 significant figures)

3. Trailing zeros are always significant for digits after the decimal point in a number, *e.g.*, 0.1200 g (4 significant figures; the two trailing zeros are significant)

4. Zeros that fall at the end of a number and after a decimal point are always significant, *e.g.*, 7.00 g (3 significant figures)

5. Zeros that fall after a non-zero digit and before the decimal point, may or may not be significant. In general, if there is no decimal point given, the zeros are not significant, *e.g.*, 310 mL (2 significant figures). If the zeros are significant, a decimal point is placed after the last zero, *e.g.*, 310. mL (3 significant figures)

Of course, in order to eliminate all confusion, it is easier to write the number using exponential notation. *All digits (including zeros) are significant for a number written in exponential notation, e.g.*, 5.01×10^{-3} mL (3 significant figures) and 5.0100×10^{-3} mL (5 significant figures).

When performing calculations, attention must be paid towards significant figures. In working problems on a calculator, *perform the calculation using all of the digits allowed by the calculator and round off only at the end of the problem.* Consider the following rules:

1. When adding/subtracting, the number of decimal places in the answer must equal the number of decimal places in the number with the *fewest* places, *e.g.*, 0.21 + 4.9 + 11.1124 = 16.2224 on a calculator, but should be reported as 16.2

2. When multiplying/dividing, the number of significant figures in the answer should be the same as that in the number with the *fewest* significant figures, *e.g.*, $15.5 \times 27.3 \times 5.4 =$ 2285.01 on a calculator, but should be reported as 2300 (or 2.3×10^3)

Keep in mind that the line of numbers which appears on your calculator's display after hitting the "=" button is a correct arithmetical calculation. For example, it would be pointless to write:

$$0.1003 \text{ g } Na_2C_2O_4 \times \left(\frac{\text{mol } Na_2C_2O_4}{134.00 \text{ g } Na_2C_2O_4} \right) = 7.708955224 \times 10^{-4} \text{ mol}$$

The answer contains far more figures that are justified. The answer should be given to 4 significant figures, specifically 7.709×10^{-4} mol. This may seem obvious, but it is a very common fault found in student calculations!

Propagation of Experimental Uncertainty

When experimental measurements are used in calculations, there is a propagation of experimental uncertainty that must be determined such that the final result of the calculation also includes the overall experimental uncertainty. *Except for standard deviation calculations* (see below), all experimental values and calculations based on them must include uncertainty limits. Sample calculations for the propagation of experimental uncertainty are given in Appendix B.

Recall that the accuracy of a measurement refers to the agreement between the measured value and the true value such that it is a measure of the correctness of the measurement. To determine accuracy, we calculate the difference between the true value (assuming it is known) and the average (or mean) value from a series of measurements. The **average** is given by:

$$\text{Average} = \bar{x} = \frac{\sum_{i=0}^{n} x_i}{n} \tag{1}$$

where Σ is the operation of summation (*i.e.*, the sum of the individual measurements), $\bar{x}$ is the average value, x_i is the individual measurements, and n is the number of measurements. The accuracy would then be defined as:

$$\text{Accuracy} = \bar{x} - \text{true value} \tag{2}$$

An accurate result is one that agrees closely with the true value (the smaller the difference, the greater the accuracy). This is sometimes called the **mean error**, and is reported in the same units as the measurement, *e.g.* if the true value = 2.62 g and $\bar{x}$ = 2.52 g, the accuracy would be -0.10 g. Many times errors are expressed in terms of a percentage. The **percent error** (sometimes called the **relative error**) is the absolute value of the difference of the true value and the experimental (measured or mean) value, divided by the true value, and multiplied by 100:

$$\text{Percent error} = \frac{\left|\text{true value} - \text{experimental value}\right|}{\text{true value}} \times 100 \tag{3}$$

Precision refers to the agreement between several measurements of the same quantity, *i.e.*, the reproducibility of a given type of measurement, and is usually described in terms of the **standard deviation**, *s*, which is defined as follows:

$$s = \sqrt{\frac{\sum_{i=1}^{n} \left(x_i - \bar{x}\right)^2}{(n-1)}} \tag{4}$$

where $(x_i - \bar{x})$ is the deviation of each measurement from the average. Actually this formula only gives an estimate of the standard deviation; an exact one would require an infinite number of measurements. Because of this, we would say that our confidence in our mean is low if we only take a few measurements.

As an example of this type of calculation, consider the following problem. In a titration experiment using a burette, you obtain the following results: 36.78 mL, 36.80 mL and 36.75 mL. You'd like to know how well these volumes agree with one another, so you proceed to calculate the standard deviation. First, calculate the mean volume:

$$\bar{x} = (36.78 + 36.80 + 36.75) \div 3$$
$$= 36.78 \text{ mL}$$

Now:

x_i (mL)	$(x_i - \bar{x})$ (mL)	$(x_i - \bar{x})^2$ (mL^2)
36.78	0.00	0.0000
36.80	0.02	0.0004
36.75	0.03	0.0009
		$\sum$ = 0.0013

So:

$$s = \sqrt{\frac{0.0013}{(3-1)}} = 0.03$$

So, the mean volume would have a standard deviation of 0.03 mL (burettes can only be read to ± 0.01 mL, so *s* is reported as 0.03 mL). We would write the answer as 36.78 mL ± 0.03 mL. The standard deviation calculation is an estimate of the probable error of a single measurement. Note that this is **not** the same as experimental uncertainty as explained above (that is, uncertainties due to inherent limitations of the instrument or user). The standard deviation of the mean is sometimes referred to as the **standard error**. Generally speaking, for a large number of measurements, the probability that a given measurement will fall within ± *s* of the average value is 68% and there is a 99.7% probability that it will fall within ± 3*s* of the average value.

Pre-Lab Assignment

1. There are four sections to this experiment (see below under **Procedure**). Describe **very briefly** (using a few sentences for each) the specific purpose behind each section (that is, what information you are trying to gain). Check the **Report** section for additional insight.

2. A student performed an analysis of a sample for its sodium content and got the following results: 18.61%, 18.58%, 18.64% and 18.59%. The actual amount (true value) of sodium in the sample is 19.34%. Calculate the mean and the standard deviation. What can you say about the accuracy and precision of these results?

3. Calculate the percent error for the following measurement. The density of an iron block determined in an experiment was 7.86 ± 0.02 g/cm^3. The true value is 8.03 ± 0.02 g/cm^3 (see *Appendix B*).

Procedure *(WORK IN PAIRS)*

• All water used in measurements may be poured down the sink

1. A copper sample is located beside each balance. Record the copper sample number, the balance number, and the temperature. Weigh the sample four times and record the masses (if you work in pairs, each student will do 2 weighings). When weighings are finished, return the copper sample to the balance number it came from.

2. Clean a 50 mL burette, place it in a stand, and add distilled water until the burette reading is about 45.00 mL. Record the reading (you should be able to read the water level to 0.01 mL). The correct procedure for reading a burette is shown in *Appendix A*. Using a 10 mL graduated cylinder, measure out 10.0 mL of distilled water and **add it to the water already in the burette**. Record the new burette reading. Measure out another 10.0 mL of water in the graduated cylinder and add it to the burette. Again record the burette reading. Add a further 10.0 mL from the cylinder and record the burette reading.

   Weigh a clean, dry weighing bottle and lid on an analytical balance. Add 10.0 mL of water **from the graduated cylinder** and re-weigh. Note the temperature of the water immediately after weighing.

3. Set up the burette as in (2), *i.e.* water level at about 45.00 mL, and repeat the procedure of (2), but this time use a 10 mL pipette to deliver 10.00 mL. The correct technique for using a pipette is shown in *Appendix A.*

   Weigh a clean, dry weighing bottle and lid on an analytical balance. Add 10.00 mL of water **from the pipette** and re-weigh. Note the temperature of the water.

4. Weigh a clean, dry weighing bottle and lid on an analytical balance. **Measure accurately about 10 mL of water from the burette** (**by taking initial and final readings on the burette**) into the weighing bottle, and then re-weigh the bottle now containing the water. Note the temperature of the water.

Report

Note: For all readings and calculations (**except for question #1, below**) include the uncertainty.

1. Report the masses of the copper, calculate its average mass, and calculate the standard deviation. (Take care with significant figures in your answers.)

2. Report **ALL** your burette readings, and determine the average volume of the 10 mL samples for each of (a) the graduated cylinder and (b) the pipette.

3. Report the initial and final masses of the weighing bottle and thus the mass of the water

for each of the three different devices. From the table of densities below, calculate the volume of the mass of water you delivered from each device. This volume calculated from the mass of water is your true value. For your experimental values, use the following: graduated cylinder, 10.0 mL; pipette, 10.00 mL; burette, use the actual volume you delivered from your burette. Calculate the percent error in your experimental result for each device.

4. List the instruments, *i.e.*, burette, graduated cylinder, and pipette, in order of **increasing accuracy**. Give your reasoning (based on percent error calculations).

Table 1: STANDARD DENSITY OF WATER†

T (°C)	Density (g/mL)									
	0.0	0.1	0.2	0.3	0.4	0.5	0.6	0.7	0.8	0.9
18	0.998598	579	560	542	523	504	485	466	446	427
19	0.998407	388	368	348	328	308	288	268	247	227
20	0.998206	186	165	144	123	102	081	059	038	016
21	0.997995	973	951	929	907	885	863	841	818	795
22	0.997773	750	728	705	682	658	635	612	588	562
23	0.997541	517	494	470	446	422	397	373	349	324
24	0.997299	275	250	225	200	175	150	124	099	074
25	0.997048	022	997*	971*	945*	919*	893*	866*	840*	814*
26	0.996787	760	734	707	680	653	626	599	571	544
27	0.996517	489	461	434	406	378	350	322	294	266

† Source: "Handbook of Chemistry and Physics" 74th ed., D.R. Lide ed. Ann Arbor, MI: CRC Press.

As an example of reading this table, at 23°C, the density is the number in the 0.0 column, 0.997541 g/mL. To determine the density at 23.7°C, you replace the last three digits from the 0.0 column with the three digits from the 0.7 column, giving you a density of 0.997373 g/mL.

APPENDIX A

WEIGHING AND VOLUMETRIC EQUIPMENT AND TECHNIQUES

Analytical Balance

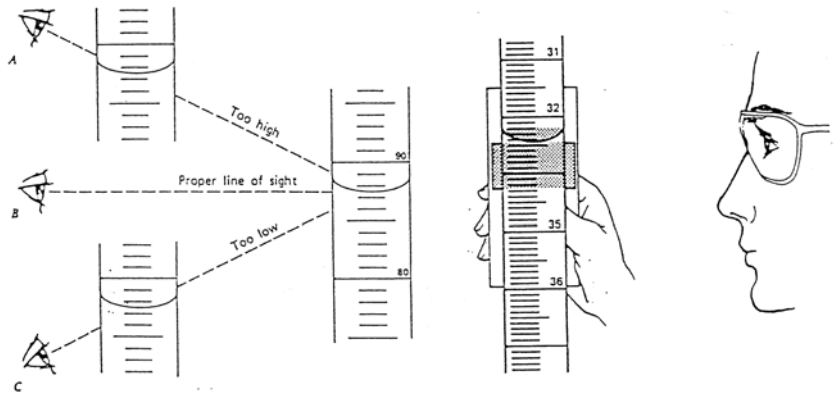
The analytical balances in the lab can be read to four decimal places (*i.e.,* to the nearest 0.0001 g or 0.1 mg), with uncertainties of ±0.0001 g. **All analytical weighings are made by difference** (never add your reagent to the weighing container while it is on the pan), so the effect of a possible systematic error in the balance (*e.g.,* always providing readings 0.01 g higher) is not important, **provided that the two readings are made on the same balance!** Be sure not to rest your arm or apply any pressure on the balance table, as the balance is quite sensitive to slight changes in position. When performing a weighing or zeroing the balance, make sure the sliding doors are closed, as even small wind currents will frustrate balance readings. Make sure you are aware of the maximum allowable mass the balance can sustain before placing your weighing container (*e.g.,* boat, small beaker, etc.) on the pan, or measuring out a specific quantity. Most four-place balances cannot be used for masses greater than about 150 to 250 g. In the event that some of your reagent is spilled onto the pan or within the weighing chamber of the balance, **clean it up immediately.**

Volumetric Equipment and Techniques

A. Reading a Meniscus. For exacting measurements of clear or transparent liquids in graduated cylinders, pipettes, burettes and volumetric flasks, the volume of a solution is read at the bottom of the meniscus. Steady the eye horizontal to the surface of the liquid [Figure A1(a)]; position the top edge of a black mark (made on a white card) just below the meniscus. The black reflection in the meniscus, viewed against the white card, precisely defines the liquid level [Figure A1(b)]. Substituting a finger for the black mark on a white card is not as effective, but it does help.

B. Cleaning of Volumetric Equipment. Wash the glassware with a small amount (it is rather concentrated) of the detergent solution. There are brushes available in the lab designed to wash the various pieces of equipment. Rinse the glassware with tap water and then finally with distilled water. No water droplets will adhere to the inner walls of the glassware if it is cleaned properly.

C. Volumetric Flasks. Volumetric flasks are pear-shaped, to facilitate drainage, and have a narrow neck to increase the precision of the volume reading at the single graduation mark. They are calibrated to contain a specified volume. A weighed amount (or measured volume) of substance is added to the flask, and solvent is added precisely to the graduation mark. The flask is capped and inverted a number of times to ensure thorough mixing. The temperature of the solution should be close to the calibrated value (a 5°C variation will cause an error of about 1 mL in 1000 mL – that is a 0.02% error per degree).



(a) (b)

Figure A1. (a) Reading the meniscus without parallax error. (b) Reading the meniscus using a "meniscus reader".

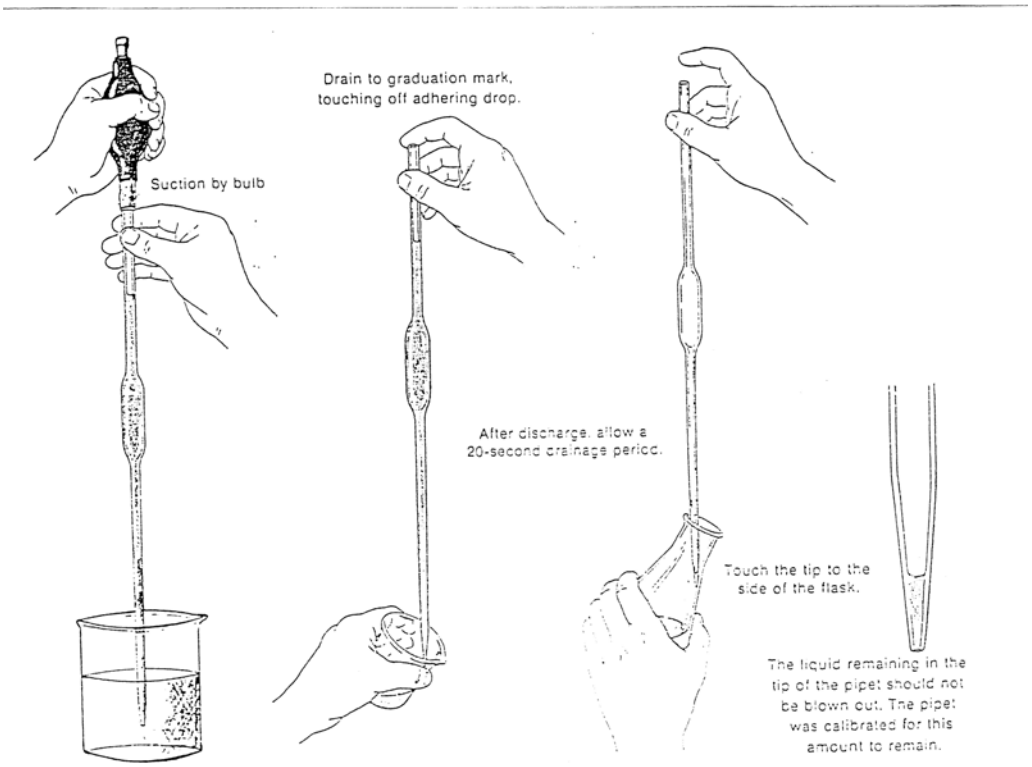


Figure A2. Technique for measuring liquids with a pipette (for a

right-handed person).

D. Pipettes. Transfer pipettes are designed to deliver a single, fixed volume of liquid with good precision. Transfer the solution that you intend to pipette from the reagent bottle to a beaker. To avoid contamination, **never** pipette directly from a reagent bottle. Dry the tip of the pipette with a tissue. Using suction from a rubber bulb (**Caution: NEVER use your mouth**), draw 2-3 small portions into the pipette as a rinse (Figure A2). Roll each rinse around in the pipette to make sure the solution washes the entire inner surface. Deliver each rinse to a waste beaker.

Use the suction bulb to fill the pipette so that the liquid level is above the graduation mark (but not into the bulb!). Remove the bulb and quickly put your index finger over the end of the pipette. Roll your finger over to the side (so that air can enter the pipette slowly) and allow the liquid to drain until the bottom of the meniscus is exactly on the graduation mark.

The tip of a pipette should be small enough to require 20 seconds to empty a 10 mL pipette, 30 seconds for a 50 mL pipette, and so forth. If the outflow rate is faster, this much time should be allowed for drainage. The pipette is maintained in a vertical position during discharge, with the contacting wall of the receiving vessel at about 45º. Finally, the adhering drop at the tip is touched off inside the receiving vessel, but the small portion of liquid in the tip is left in place.

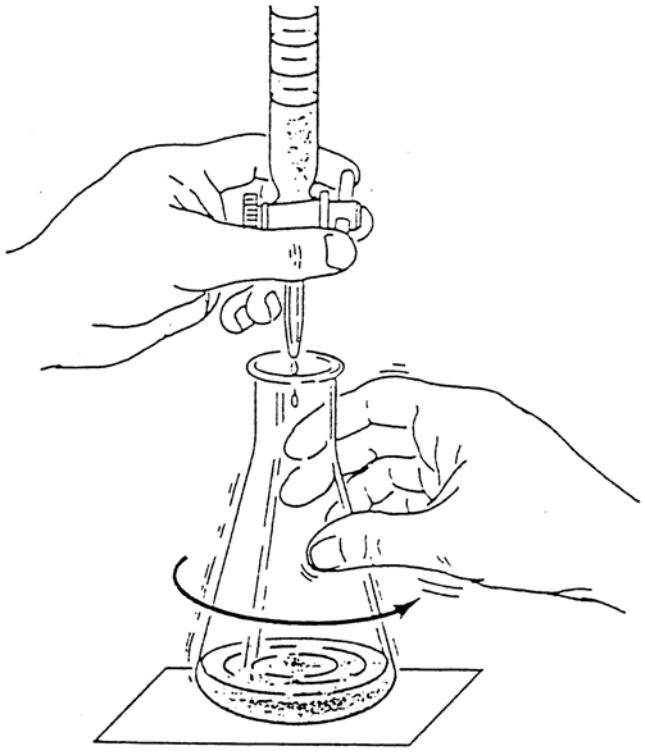


Figure A3. Technique for a titration by a right-handed person.

E. Burettes. Burettes are long cylindrical tubes of very uniform bore, calibrated along their length, with some device to control the flow of liquid - a glass stopcock or a Teflon valve. Burettes are calibrated to deliver, not to contain, the measured volume of liquid. If you are using a particular burette for the first time, examine its markings before you fill it. The lines that span the entire

circumference occur for each mL, starting with zero at the top and reaching the maximum volume at the bottom. As a consequence, the burette will show the volume of a liquid that has been delivered rather than the volume that remains. On a 50 mL burette, the smaller lines indicate each tenth of a mL. The spacings between these lines will allow you to estimate the volume **to the nearest 0.01 mL**. Thus typical burette readings would be 9.34 mL and 17.60 mL. Readings such as 9.3 mL and 17.6 mL are not acceptable.

As with a pipette, the burette must be rinsed with small amounts of the solution to be measured. Then fill the burette to above the zero mark with the stopcock closed. Open the stopcock fully once or twice so that the liquid drains rapidly to flush out any air bubbles in the tip of the burette. Drain the burette until the meniscus rests between the zero and the 1 mL mark. Do not waste time trying to align the bottom of the meniscus with the zero or any other specific mark. Read the burette with your eye on the same level as the meniscus. To obtain the volume of the liquid that you delivered, subtract this reading from the final reading.

The tip must be small enough that the delivery time is sufficient to allow for drainage (about 40 seconds for 25 mL or 90 seconds for the full 50 mL). If the outflow rate is faster, this much time must be allowed before reading the final volume. Figure A3 shows the best technique for a titration by a right-handed student. Note that the left hand is used to open and close the stopcock. With a bit of practice, you will be able to adjust the stopcock so that as little as half a drop will form on the tip. This drop then is touched off onto the inside neck of the receiving flask and washed down with distilled water (unless conducting a potentiometric titration, *i.e.*, using a pH meter). The right hand is used to swirl the flask as the titrant is added. A left-handed student would turn the burette 180º, control the stopcock with the right hand and swirl with the left hand. Coloured endpoints are most easily seen if a white background, such as a white tile, is used. As the endpoint is neared the colour of the indicator persists longer in the solution and the addition of titrant is slowed down. Note that, as shown in Figure A3, the tip of the burette is kept close to the mouth of the flask to avoid errors due to splattering. Experienced chemists usually perform a quick, preliminary titration. Although the endpoint may be overshot, it gives the chemist an idea of the size of the titre to be expected and is preferable to doing the first titration very slowly and carefully when the size of the titre is unknown. This procedure, of course, depends on the amount of solution that is available for analysis.

APPENDIX B

THE PROPAGATION OF EXPERIMENTAL UNCERTAINTY IN CALCULATIONS

Significant Figures and Calculations

For this course we will be using the definitions and rules for significant figures that are given in the textbook Chemistry: the Molecular Science by Moore *et al*. in Appendix A. Please note that the rules for addition or subtraction are different than for multiplication or division.

Experimental Uncertainty

There are several methods for determining overall uncertainty when measurements are combined in calculations. Below are simple procedures for estimating the maximum experimental uncertainty. (Although easy to use, these calculations slightly overestimate the experimental uncertainty.)

Addition and Subtraction. When measured quantities are added or subtracted, the maximum absolute uncertainty in the result is the sum of the uncertainty in each measurement.

i.e.: **$(A \pm a) + (B \pm b) = (C \pm c)$, where $C = A \pm B$ and $c = a + b$**

For example, in Experiment #1 you may obtain data such as these:

mass of weighing bottle:	12.2835 ± 0.0001 g
mass of bottle + water:	22.2708 ± 0.0001 g
mass of water:	$9.9873 \pm$ **0.0002** g

This result can be verified by using data which yield the largest and smallest mass of water:

12.2834 g		12.2836 g
22.2709 g	and	22.2707 g
9.9875 g		9.9871 g

The mean value is indeed 9.9873, with uncertainty $\pm$ 0.0002 g.

Multiplication and Division. When measured quantities undergo the operations of multiplication or division, the derived result has a maximum relative (*i.e.*, fractional or percentage) uncertainty which is the sum of the relative uncertainties in each measurement.

i.e. **$(A \pm a) \times (B \pm b) = (C \pm c)$, where $C = A \times B$ and $c = C \times (a/A + b/B)$**
or **$(A \pm a) \div (B \pm b) = (C \pm c)$, where $C = A \div B$ and $c = C \times (a/A + b/B)$**

For example, what is the mass of a 20.2 ± 0.1 mL of mercury which has density, at 20°C, of 13.4562 ± 0.0001 g mL^{-1}?

The sum of the relative uncertainties (a/B + b/B) is:

$$(0.1/20.2) + (0.0001/13.5462) = (5.0 \times 10^{-3}) + (7.4 \times 10^{-6}) = 5.0 \times 10^{-3}$$

In this case, the relative uncertainty in the density is negligibly small. The result for the overall calculation is:

mass = density × volume

$= [(13.5462 \text{ g mL}^{-1}) \times (20.2 \text{ mL})] \pm [(13.5462 \times 20.2) \times (5.0 \times 10^{-3})]$

$= 274 \pm 1$ g (with correct significant figures)

or, with one extra significant figure,* $273._6 \pm 1._4$ g.

Logarithms. When a measured quantity undergoes a logarithmic operation, the absolute uncertainty of the derived result is the relative uncertainty of the measurement. The general formula for any base logarithm is:

$$\textbf{logarithm}_x(A \pm a) = \log_x(A) \pm (a/A \ln x), \text{ where } \ln = \log_e$$

For example, if the rate constant **k** of a chemical reaction is determined to be 1240 ± 50 s^{-1}, the natural (base-e) logarithm (ln) is:

$\ln \mathbf{k} = \ln(1240) \pm [50/(1240 \times \ln e)]$

$= \ln(1240) \pm [50/(1240 \times 1)]$

$= 7.12_2 \pm 0.04_0$

* An extra significant figure (often subscripted) is used to eliminate rounding-off errors in subsequent calculations.

If the base-10 logarithm (log) is required, then the calculation becomes:

$\log_{10}$ **k** = $\log_{10}(1240) \pm [50/(1240 \times \ln 10)]$

= $\log_{10}(1240) \pm [50/(1240 \times 2.303)]$

= $3.09_3 \pm 0.01_7$

The opposite relationship holds for inverse logarithms. Consider as an example: what is the value of $e^{-\Delta H/(RT)}$, where ΔH = 48.7 ± 0.1 kJ mol^{-1}, T = 293.2 ± 0.2 K and the value R is the constant 8.31451 J $K^{-1}mol^{-1}$?

$-\Delta H/(RT)$ = $-[48.7 \times 10^3/(8.31451 \times 293.2)]$

= -19.977

note: 48.7 kJ was converted to 48.7×10^3 J because of the units in R.

The uncertainty = $19.977 \times [(0.1/48.7) + (0.2/293.2)] = 0.055$

So, $e^{-\Delta H/(RT)}$ = $e^{-19.977} \pm [(e^{-19.977}) \times 0.055]$

= $(2.1_1 \pm 0.1_2) \times 10^{-9}$

APPENDIX C

GRAPHING EXPERIMENTAL DATA

Guidelines

1. **Graph paper.** Use quality graph paper, with mm graduations, for constructing hand-drawn plots. Computer-drawn plots are acceptable, provided that *all* the guidelines are followed. Most graphing and spreadsheet packages include statistical software that provide quick and convenient analysis of linear and non-linear data correlations, along with the relevant experimental uncertainty.

2. **Titles.** All figures, graphs included, must be clearly titled. For example:
   Figure 3. Student lab grades correlated with student height.
   All labelling (titles, axes, etc.) should be in ink!

3. **Axes.** Axes should be clearly drawn on the graph, along with "ticks" (pointing inwards) indicating the positions that correspond to the numerical labels. Each axis should have a title that includes variable name and units. For example:
   Student Height (cm)

4. **Scale.** Ensure that the graph comfortably fits the entire page. It is not necessary that the axes start at 0.

5. **Data points.** Data points should be highlighted by the use of appropriate symbols. For example: ○ □ ◇ △ ▽ are all acceptable. Only place two or more data sets on a single graph if they are reasonably separated and a symbol key is provided.

6. **Correlations.** Never connect data point to data point. Your line/curve should convey the *overall* relationship between the variables. If the relationship seems to be linear, us a ruler to draw the best line fit (with points scattered equally above and below the line). Nonlinear correlations should be indicated with a neatly drawn curve. To calculate the slope of a straight line, choose two well separated points that lie directly on the line (NOT ACTUAL DATA POINTS) and divide the *rise* (difference between the *y* coordinates) by the *run* (difference in *x* coordinates). The experimental uncertainty in the slope can be estimated as the difference in slope between the best fit and the steepest/shallowest lines that can reasonably be drawn.