2I THE LABORATORY NOTEBOOK

A laboratory notebook is needed to record measurements and observations concerning an analysis. The book should be permanently bound with consecutively numbered pages (if necessary, the pages should be hand numbered before any entries are made). Most notebooks have more than ample room, so there is no need to crowd entries.

The first few pages should be saved for a table of contents that is updated as entries are made.

2I-1 Maintaining a Laboratory Notebook

1. *Record all data and observations directly into the notebook in ink.* Neatness is desirable, but you should not achieve neatness by transcribing data from a sheet of paper to the notebook or from one notebook to another. The risk of misplacing—or incorrectly transcribing—crucial data and thereby ruining an experiment is unacceptable.
2. Supply each entry or series of entries with a heading or label. A series of weighing data for a set of empty crucibles, for example, should carry the heading "empty crucible mass" (or something similar), and the mass of each crucible should be identified by the same number or letter used to label the crucible.
3. Date each page of the notebook as it is used.
4. *Never* attempt to erase or obliterate an incorrect entry. Instead, cross it out with a single horizontal line and locate the correct entry as nearby as possible. Do not write over incorrect numbers. With time, it may become impossible to distinguish the correct entry from the incorrect one.

Remember that you can discard an experimental measurement *only if you have certain knowledge that you made an experimental error.* Thus, you must carefully record experimental observations in your notebook as soon as they occur.

An entry in a laboratory notebook should never be erased but should be crossed out instead.

2J SAFETY IN THE LABORATORY

There is necessarily a degree of risk associated with any work in a chemical laboratory. Accidents can and do happen. Strict adherence to the following rules will go far toward preventing (or minimizing the effect of) accidents:

1. Before you begin work in any laboratory, learn the location of the nearest eye fountain, fire blanket, shower, and fire extinguisher. Learn the proper use of each, and do not hesitate to use this equipment if the need arises.
2. *Wear eye protection at all times.* The potential for serious and perhaps permanent eye injury makes it mandatory that adequate eye protection be worn at all times by students, instructors, and visitors. Eye protection should be donned before entering the laboratory and should be used continuously until it is time to leave. Serious eye injuries have occurred to people performing such innocuous tasks as computing or writing in a laboratory notebook. Incidents such as these usually result from someone else's loss of control over an experiment. Regular prescription glasses are not adequate substitutes for eye protection approved by the Office of Safety and Health Administration (OSHA). Contact lenses should never be used in the laboratory because laboratory fumes may react with them and have a harmful effect on the eyes.

3. Most of the chemicals in a laboratory are toxic, some are very toxic, and some—such as concentrated solutions of acids and bases—are highly corrosive. Avoid contact between these liquids and the skin. In the event of such contact, *immediately* flood the affected area with large quantities of water. If a corrosive solution is spilled on clothing, remove the garment immediately. Time is of the essence, so do not be concerned about modesty.
4. *NEVER* perform an unauthorized experiment. Unauthorized experiments are grounds for disqualification at many institutions.
5. Never work alone in the laboratory. Always be certain that someone is within earshot.
6. Never bring food or beverages into the laboratory. *NEVER* drink from laboratory glassware. *NEVER* smoke in the laboratory.
7. Always use a bulb or other device to draw liquids into a pipet. *NEVER* pipet by mouth.
8. Wear adequate foot covering (no sandals). Confine long hair with a net. A laboratory coat or apron will provide some protection and may be required.
9. Be extremely tentative in touching objects that have been heated because hot glass looks exactly like cold glass.
10. Always fire-polish the ends of freshly cut glass tubing. *NEVER* attempt to force glass tubing through the hole of a stopper. Instead, make sure that both tubing and hole are wet with soapy water. Protect your hands with several layers of towel while inserting glass into a stopper.
11. Use fume hoods whenever toxic or noxious gases are likely to be evolved. Be cautious in testing for odors. Use your hand to waft vapors above containers toward your nose.
12. Notify your instructor immediately in the event of an injury.
13. Dispose of solutions and chemicals as instructed. It is illegal to flush solutions containing heavy metal ions or organic liquids down the drain in most localities. Alternative arrangements are required for the disposal of such liquids.

Reference

[9]See also Howard M. Kanare, *Writing the Laboratory Notebook*, Washington, DC: American Chemical Society, 1985.

4B-2 Density and Specific Gravity of Solutions

Density and specific gravity are related terms often found in the analytical literature. The **density** of a substance is its mass per unit volume, and its **specific gravity** is the ratio of its mass to the mass of an equal volume of water at 4°C. Density has units of kilograms per liter or grams per milliliter in the metric system. Specific gravity is dimensionless and so is not tied to any particular system of units. For this reason, specific gravity is widely used in describing items of commerce ~~(see Figure 4-1)~~. Since the density of water is approximately 1.00 g/mL and since we use the metric system throughout this text, we use density and specific gravity interchangeably. The specific gravities of some concentrated acids and bases are given in ~~Table 4-3~~.

Density expresses the mass of a substance per unit volume. In SI units, density is expressed in units of kg/L or alternatively g/mL.

Specific gravity is the ratio of the mass of a substance to the mass of an equal volume of water.

EXAMPLE 4-10

Calculate the molar concentration of HNO_3 (63.0 g/mol) in a solution that has a specific gravity of 1.42 and is 70.5% HNO_3 (w/w).

Solution

Let us first calculate the mass of acid per liter of concentrated solution

$$\frac{\text{g HNO}_3}{\text{L reagent}} = \frac{1.42\ \cancel{\text{kg reagent}}}{\text{L reagent}} \times \frac{10^3\ \cancel{\text{g reagent}}}{\cancel{\text{kg reagent}}} \times \frac{70.5\ \text{g HNO}_3}{100\ \cancel{\text{g reagent}}} = \frac{1001\ \text{g HNO}_3}{\text{L reagent}}$$

Then,

$$c_{\text{HNO}_3} = \frac{1001\ \cancel{\text{g HNO}_3}}{\text{L reagent}} \times \frac{1\ \text{mol HNO}_3}{63.0\ \cancel{\text{g HNO}_3}} = \frac{15.9\ \text{mol HNO}_3}{\text{L reagent}} \approx 16\ \text{M}$$

Figure 4-1 Label from a bottle of reagent-grade hydrochloric acid. Note that the specific gravity of the acid over the temperature range of 60° to 80°F is specified on the label. (*Label provided by Mallinckrodt Baker, Inc., Phillipsburg, NJ 08865*)

TABLE 4-3

Specific Gravities of Commercial Concentrated Acids and Bases

Reagent	Concentration, % (w/w)	Specific Gravity
Acetic acid	99.7	1.05
Ammonia	29.0	0.90
Hydrochloric acid	37.2	1.19
Hydrofluoric acid	49.5	1.15
Nitric acid	70.5	1.42
Perchloric acid	71.0	1.67
Phosphoric acid	86.0	1.71
Sulfuric acid	96.5	1.84

EXAMPLE 4-11

Describe the preparation of 100 mL of 6.0 M HCl from a concentrated solution that has a specific gravity of 1.18 and is 37% (w/w) HCl (36.5 g/mol).

Solution

Proceeding as in Example 4-10, we first calculate the molar concentration of the concentrated reagent. We then calculate the number of moles of acid that we need for the diluted solution. Finally, we divide the second figure by the first to obtain the volume of concentrated acid required. Thus, to obtain the concentration of the reagent, we write

$$c_{\mathrm{HCl}} = \frac{1.18 \times 10^3 \cancel{\text{g reagent}}}{\text{L reagent}} \times \frac{37 \cancel{\text{g HCl}}}{100 \cancel{\text{g reagent}}} \times \frac{1 \text{ mol HCl}}{36.5 \cancel{\text{g HCl}}} = 12.0 \text{ M}$$

The number of moles HCl required is given by

$$\text{no. mol HCl} = 100 \cancel{\text{mL}} \times \frac{1 \cancel{\text{L}}}{1000 \cancel{\text{mL}}} \times \frac{6.0 \text{ mol HCl}}{\cancel{\text{L}}} = 0.600 \text{ mol HCl}$$

Finally, to obtain the volume of concentrated reagent, we write

$$\text{vol concd reagent} = 0.600 \cancel{\text{mol HCl}} \times \frac{1 \text{ L reagent}}{12.0 \cancel{\text{mol HCl}}} = 0.0500 \text{ L or } 50.0 \text{ mL}$$

Therefore, dilute 50 mL of the concentrated reagent to 600 mL.

The solution to Example 4-11 is based on the following useful relationship, which we will be using countless times:

$$\boxed{V_{\mathrm{concd}} \times c_{\mathrm{concd}} = V_{\mathrm{dil}} \times c_{\mathrm{dil}}} \qquad (4\text{-}4)$$

where the two terms on the left are the volume and molar concentration of a concentrated solution that is being used to prepare a diluted solution having the volume and concentration given by the corresponding terms on the right. This equation is based on the fact that the number of moles of solute in the diluted solution must equal the number of moles in the concentrated reagent. Note that the volumes can be in milliliters or liters as long as the same units are used for both solutions.