SAFETY GUIDELINES

Teaching science using student activities and teacher demonstrations has the potential for danger. Thus it is important to carefully plan student activities or teacher demonstrations to ensure that safety is paramount. The following are some guidelines to follow in carrying out student activities safely:

A. General

1. Student work surfaces should be level; several flat-topped tables may be placed together to provide a larger work surface for students working in groups. Desks with sloped surfaces are not suitable.
2. Eye protection is needed any time there is the possibility of liquids or solids being splashed or propelled into the eyes.
3. Do not tolerate improper student behaviour (e.g. fooling around) during student activities.
4. If students spill chemicals on their skin or clothing, the affected area should be washed immediately with soap and water.
5. Students should wash their hands at the end of any activity involving the use of chemicals.
6. Students should never taste chemicals, unless specifically asked to do so.
7. Work with the smallest volumes and quantities of reagents possible. Never exceed the recommended quantities.
8. When assembling clamps, etc. to a stand, ensure that they are firmly attached. Similarly, ensure that a clamp used to support a test tube, for example, is tight enough to prevent the test tube from falling.
9. Provide students with solutions in plastic dropper bottles and solid (powdered) chemicals in small vials to avoid spillage and contamination of the stock bottles. Label each container properly.
10. Do not use mercury thermometers. Do not place thermometers near the each of a table, where there is the danger of them rolling off.
11. If running water and a sink are not available in the classroom, provide a supply of water in a large container as well as a container for disposal of liquid wastes.
12. Rinse all glassware with clean water immediately after use to prevent chemicals from drying on the glass.

B. Heating

1. If an alcohol burner is used, ensure that it is not placed near the edge of a table. The burner should be lit only when doing an actual experiment.
2. Never leave a lit burner or candle unattended.
3. Long hair and loose clothing must be tied back to avoid accidentally coming in contact with an open flame.
4. When heating a liquid in a test tube, hold the test tube with a test tube clamp. Move the test tube back and forth in the flame so that the liquid does not superheat in one region of the test tube. Point the open end of the test tube away from other people in case the liquid shoots out.
5. Using only heat-resistant glassware (e.g. Pyrex™ or Kimax™) when heating liquids. Do not use cracked glassware.
6. Heated materials should be allowed to cool before handling, unless a suitable clamp or pair of tongs is used to hold the material.
7. Keep flammable materials away from open flames.
8. Know where the nearest fire extinguisher is located and how to operate it.

C. Electrical

1. Ensure that all electrical equipment is in good condition (e.g. no frayed or bare wires, proper grounding).
2. If extension cords are required, place them so that students will not trip over them and/or tape them to the floor. Do not overload an extension cord.

D. Chemicals

1. Generally purchase only small amounts of chemicals.
2. Store all chemicals in a lockable cabinet to limit access to the chemical supplies.
3. Label all vials and bottles used for dispensing or storing reagents and solutions clearly.
4. When diluting a concentrated acid, always add the acid to water, not *vice versa*.
5. Keep all containers tightly closed.
6. Know which chemicals are not permitted in the school.
7. Dispose of waste chemicals properly.

PREPARATION OF COMMON SOLUTIONS

For many demonstrations and experiments, it is convenient to have solutions that are prepared ahead of time and stored in dropper bottles, so that the teacher can just take them from storage and set them out for student use or for a demonstration. Dropper bottles are available from science supply companies – the 60 mL size is recommended. For larger quantities of solution, glass or plastic juice bottles may be used. BE SURE TO LABEL EACH SOLUTION WITH NAME, CONCENTRATION, AND DATE OF PREPARATION.

Sources of Water

Generally tap water should not be used for preparing solutions, since it may contain dissolved substances that may interfere with the reagents you are adding to the water. There are several sources of purified water that you could use.

a) Distilled water – many grocery stores carry jugs of distilled water at a relatively low cost. Or perhaps someone in your community has water still and would be willing to supply you with distilled water.

b) Reverse Osmosis (RO) water – the water is forced under pressure through a membrane permeable only to water, leaving the minerals behind.

c) Deionised (DI) Water – normal tap water is passed through a special unit which removes all the dissolved minerals in the water, but not the organic matter. A normal water softener is not sufficient, as it does not remove all the minerals (it replaces the hardness ions, calcium and magnesium, with sodium ions).

d) Condensed water from an air conditioner or dehumidifier – this water contains very low levels of dissolved minerals, but may contain dust particles. Letting the water stand for a while or filtering the water may remove most of the solid dust particles.

Concentration of Solutions

Several different methods are used to express the concentration of a solution:

a) Mass/volume

A 1% solution of sodium hydroxide, for example, is prepared by adding 1 g of NaOH to 100 mL of water. Similarly, a solution of 10% sodium chloride is prepared by adding 10 g of NaCl to 100 mL of water. In most cases we do not worry about the fact that the solute will increase the total volume slightly, so that you would in effect have more than 100 mL of solution.

b) Volume/volume

This is commonly used when one liquid is dissolved in another. For example, to prepare a solution of 25% methanol in water, 25 mL of methanol would be combined with 75 mL of water. Similarly, a solution of 10% hydrochloric acid is prepared by taking 10 mL of the concentrated muriatic acid (commercial hydrochloric acid) and adding it to 90 mL of water, even though muriatic acid itself contains only about 30% HCl. Please note: ***CONCENTRATED ACID IS ALWAYS ADDED TO WATER, NOT WATER TO ACID!*** The addition of water to some acids (particularly sulfuric acid) generates a lot of heat, enough to cause the water to boil!

c) **Molarity**
This is the concentration unit normally used in chemistry. Molarity refers to the number of moles of solute per litre of solution; this is simply abbreviated as M. For example, a solution containing 1.0 mole of NaCl/L would be written as 1.0 M NaCl. The mole represents a very large number (6.022×10^{23}); one mole of any substance contains the same number of molecules. A mole of a substance has a mass equal to the atomic masses of each element, expressed in grams. For example, the molar mass of water, H_2O, is 2 x 1.00 (mass of H) + 1 x 16.0 (mass of O) = 18.0 g/mole. For sodium hydroxide, NaOH, it would be 23.0 (mass of Na) + 16 (mass of O) +1.00 (mass of H) = 40 g/mole. For NaCl, the molar mass is 23.0 + 35.5 = 58.5 g/mole. (These numbers can be found on any Periodic Table.) Thus 18 g of water contains the same number of molecules as 40 g of NaOH.

To prepare a solution of 1.0 M NaCl, 58.5 g of NaCl are added to water, mixed well until dissolved, and water added until the total volume is 1.0 L. But if only 100 mL are needed, then 5.85 g of NaCl are used and water added to make 100 mL of solution. Since actual concentrations are not critical at the elementary level, it is not necessary to be overly careful in measuring the quantities of solid and water.

Preparation of Specific Solutions

(The solutions described below are stable (if properly closed) and classified as a LOW HAZARD at the indicated concentration.)

1. **Hydrochloric acid**
Muriatic acid, the common name for hydrochloric acid, is available in most hardware or builder supply stores. Its concentration is approximately 10 M (this means that 1 L of the muriatic acid contains about 365 g of HCl). To prepare 100 mL of 1 M HCl, add 10 mL of muriatic acid to 90 mL of water; this concentration is suitable for many experiments calling for hydrochloric acid.

2. **Sulfuric acid**
Drain cleaner containing sulfuric acid has a concentration of about 18 M. To prepare 1.0 L of 0.1 M H_2SO_4, add 5.6 mL of the concentrated acid to about 500 mL of water, swirl to mix well, and finally add more water to bring to a total volume of 1.0 L

3. **Sodium hydroxide** (a common base)
The purest form of sodium hydroxide (NaOH) is found in lye; NaOH is also the main ingredient in drain cleaners, such as Draino™. However, Draino™ also contains bits of aluminum. The metal may be physically separated from the pellets of NaOH to prepare the solution. To prepare 100 mL of 0.5 M NaOH, measure 2.0 g of NaOH pellets and dissolve in 100 mL of water. (Note that this would also be a 2% solution.) *Because sodium hydroxide will slowly dissolve glass, this solution should be stored in a plastic container.* Also sodium hydroxide will react with carbon dioxide in the air and so the container should be tightly closed.

4. **Ammonia** (another common base)
Household ammonia has a concentration of about 10 M (about 170 g of NH_3 in 1.0 L). To prepare 100 mL of 1.0 M NH_3, add 10 mL of the concentrated ammonia to 90 mL of water and mix well. To prepare 500 mL of 1 M NH_3, add 50 mL of the household ammonia to 450 mL of water.

5. **Limewater** (also a base)
Limewater is often used as a test for the presence of carbon dioxide. It can be prepared by adding lime (CaO) or slaked lime ($Ca(OH)_2$) (a few spoonfuls) to water in a 1 L plastic water or pop bottle. Shake well. Since calcium hydroxide is not very soluble in water, excess solid will settle to the bottom over night. The clear solution may be poured off as needed into another container (this prevents students from resuspending the undissolved solid!). To make more limewater, just add more water and lime and shake.

The following reaction occurs when carbon dioxide is bubbled through the limewater solution:

$$CO_2 + Ca(OH)_2 \longrightarrow CaCO_{3\,(solid)} + H_2O$$

The calcium carbonate ($CaCO_3$) that is formed is insoluble in water and turns the limewater milky or cloudy. To show students that they exhale CO_2, have them blow gently through a limewater solution in a test tube (about one third full) using a straw.

6. **Preparation of phenolphthalein solution**
Dissolve about 0.05 g of phenolphthalein powder in 50 mL of methanol or isopropyl alcohol. Then add 50 mL of water. Store in a dropper bottle; this solution is stable for years. Alternatively, if you have the older Ex Lax™ tablets available, crush about ¼ tablet with methanol in a test tube, allow the solids to settle to the bottom and draw off the clear solution using a pipette. The solution may be diluted with methanol and water and stored in a dropper bottle. Actual concentration is not critical.

7. **Preparation of other solutions from solid reagents:**
The chart below lists the amount of solute required to prepare 100 mL of 0.5 M solutions. (Many chemicals exist as hydrates, which have water molecules loosely bound to the crystals of the chemical; this is designated by the symbol "$\cdot xH_2O$".) A concentration of 0.5 M will generally give good results for tests done with solutions. In most cases, actual concentrations are not critical, so these numbers can be used as a general guide.

Reagent	**Mass needed for 100 mL of 0.5 M solution**	**Notes**
Calcium Chloride, $CaCl_2$	5.5 g	
Copper sulfate, $CuSO_4\cdot 5H_2O$	12.5 g of bluestone	Nice blue solution.
Magnesium sulfate, $MgSO_4\cdot 7H_2O$	12.3 g of Epsom salts	
Potassium nitrate, KNO_3	5.1 g of saltpeter	
Sodium bicarbonate, $NaHCO_3$	4.2 g of baking soda	
Sodium carbonate, $Na_2CO_3\cdot 10H_2O$	14.3 g of washing soda	
Sodium chloride, NaCl	2.9 g of salt	Use pickling salt, not table salt
Sodium nitrate, $NaNO_3$	4.3 g of nitrate of soda	