Teaching Guide:
Scaling a flashy demonstration

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The purpose of this presentation is to assist instructors in the selection and modification of demonstrations to account for the physical and safety limitations of the instructional setting. The classic flame test demonstration will be used to illustrate how a demonstration can be scaled up or down as dictated by such limitations. For small classrooms, Q-tips® dipped in aqueous solutions of metal salts can be placed in a candle, burner, or alcohol lamp flame [3, 4, 6]; for medium sized rooms methanol solutions of the salts in spray bottles can be sprayed into the flame of a meeker burner [2]. And, for very large lecture halls and for outdoor performances, one can ignite balloons containing water/methanol mixtures of the salts and hydrogen gas [1].

This paper addresses the following concepts:

- Atomic emission
- Atomic structure
- Light as a part of the electromagnetic spectrum
- Energy, and the relationship between color, wavelength, and energy
- Qualitative and quantitative measurements
- Prediction and trend identification
- Safety, and consideration of setting and conditions

Three Versions of the Flame Test Demonstration

Personal Protection:

Lab Coat, safety goggles or glasses, and protective gloves

Part A—Flame Test Demonstration for Small Classrooms

Materials:

- Salts: LiCl, $CuCl_2$, KCl or CsCl, $SrCl_2$, $CaCl_2$, NaCl and boric acid
- ~700 mL H_20 (de-ionized)
- Seven 100-mL beakers
- Balance
- Stirring rods
- Cotton swabs (i.e. Q-tips®)
- Bunsen burner, candle, or alcohol lamp
- Fire extinguisher
- Matches or lighter
- Seven bottles (100-mL capacity or greater) with stoppers
- Label tape and marker
- Cotton gloves, chemical gloves, goggles

Advanced Preparation:

Prepare 1 M solutions (~100 mL) of each of the salt samples (the salt may not completely dissolve due to saturation). Place each solution in separate sealed bottles. Label each bottle.

Performing the Demonstration:

Place 10-50 ml of each solution in separate 100ml beakers (label each beaker) and place a few cotton swabs by each beaker. Light the candle, burner, or alcohol lamp. Dip one end of a swab in selected aqueous solution. Place the dipped end of the swab into the flame and have students observe the color produced. Repeat procedure for each solution using a new swab. The general rule to follow is this; the stronger your emission flame, the less flammable your salt solution. If you use a candle, or an alcohol lamp, you need a methanol and salt solution in order to emit the desired color. Have a large beaker of water in which to drop the alcohol soaked cotton swabs. If you have a natural gas, or propane burner, use saturated aqueous solutions, or a salt, water and methanol mixture. The cotton swabs emit light, but do not ignite unless completely dry. I generally do not use ethanol or isopropanol for the salt solutions, although they may be used as substitutes if necessary.

Hazards:

All solutions except NaCl(aq) are ingestion hazards. $CuCl_2$(s), Boric Acid(s) and $SrCl_2$(s) are inhalation hazards. They are of mild to moderate toxicity. Fire hazards due to open flames are present. Be sure all flammable materials are removed from the area. Alcohol lamps can be extremely hazardous due to the flame being nearly colorless. Ensure water is present at every station for quenching the cotton swabs. Have a waste basket sized container set up to collect all the used cotton swabs as chemical waste. There may be quite a few. One may substitute Erlenmeyer flasks for beakers if there is concern about the chance the containers get handled carelessly. Fill the beakers and flasks used less than half full at all times, and I recommend a quarter full, refreshing the containers after each class period of use.

Disposal:

$SrCl_2$, $CaCl_2$, and $CuCl_2$ solutions should be evaporated and added to solid metal waste containers for disposal. Other solutions may be flushed down the drain with excess water.

Part B—Flame Test Demonstration for Medium Sized Classrooms

Materials:

- LiCl(s), $CuCl_2$(s), KCl(s) or CsCl(s), $SrCl_2$(s), $CaCl_2$(s), Boric Acid(s), NaCl(s)
- ~1.4 L methanol
- Meeker burner
- Fire extinguisher
- Matches or lighter
- Seven 6-oz. spray bottles
- Flame retardant board

Advanced Preparation:

Make saturated solutions (~200 mL) of each salt in methanol. Filter, if necessary, or decant solutions into spray bottles. Label each bottle. Set spray bottles to spray a mist, rather than a stream.

Performing the Demonstration:

Place burner on flame retardant board. Light the burner and adjust flame to maximum intensity. Spray solution into burner flame in such a way that the spray output is parallel to the audience. Have students observe the color produced. Repeat procedure for each solution.

Hazards:

All solutions except NaCl (aq) are ingestion hazards. $CuCl_2$(s), Boric Acid(s) and $SrCl_2$(s) are inhalation hazards. They are of mild to moderate toxicity. Methanol is toxic by ingestion or inhalation. Be sure room has adequate ventilation. All methanol solutions are highly flammable. Place containers not in use at least 1 meter away from open flame. Do not use metal salts of oxides, nitrates, superoxides, peroxides, or other anions.

Disposal:

Solutions should be placed in a designated organic waste container.

Part C—Flame Test Demonstration in Large Lecture Halls and for Outdoor Performances

Materials:

- LiCl(s), $CuCl_2$(s), KCl(s) or CsCl(s), SrCl2(s), CaCl2(s), Boric Acid(s), NaCl(s)
- Translucent helium quality balloons
- Hydrogen gas
- Water
- Methanol

- At least 7 droppers
- Mortar and pestle
- Seven small (50-100 mL) beakers
- Matches or lighter
- Flame wand, or candle on meter stick
- Scissors
- String
- 7/8" Hex nuts, or other keeper weights
- Ear protection

Advanced Preparation

Add ~1ml of water and ~1ml of methanol to ~2g of solid metal salt. Mix to make a slurry of the salt in the H_2O/MeOH mixture. Draw ~1ml into clean dropper and place solution in balloon. Inflate the balloon with hydrogen gas (volume depends upon the capacity of the balloon). Tie off balloon and attach one end of a string to the balloon and the other end to a weight (hex nut). Each balloon should float at a minimum distance of 1 meter above the surface.

Performing the Demonstration:

Select one balloon to be ignited. Keep the remaining balloons away from the area where you will ignite the balloon. Check to be sure there is plenty of air space surrounding the balloon on all sides including above it. Wear ear protection and warn the students to cover their ears because an explosion will occur (no surprises!). With a lighted candle attached to a long stick or a flame wand ignite the balloon. Have students observe the color of the flame produced. Repeat procedure with each balloon assembly.

Hazards:

All solutions except NaCl(aq) are ingestion hazards. $CuCl_2$(s), Boric Acid(s) and $SrCl_2$(s) are inhalation hazards. They are of mild to moderate toxicity. Methanol is toxic by ingestion or inhalation. Be sure room has adequate ventilation. All methanol solutions are highly flammable. Place containers not in use at least 1meter away from any open flame. Do not use metal salts of oxides, nitrates, superoxides, peroxides, or other anions.

Disposal:

Solutions should be placed in a designated organic waste container.

Discussion:

Factors Considered in the Determination of which Version to Employ:

Which version of the flame test demonstration to use is based on the needs of the teacher and students, the size and construction of the hall, and the setting in which the teaching activity is conducted.

The needs of the teacher and students refer to what kind of activity will get the point across for your students. This includes the age and number of students, aspect of the chemistry in focus for the lesson, time allotment, and level of expectation for independent or interactive study. For example, a lab activity setting with students in elementary school may view the activity as a demo in a small, intimate, least intimidating setting, using Q-tips®, and a burner, or as a middle school lab activity, with students working in groups recording their observations and comparing the different results for different elemental chloride salts. The students can observe to colors, compare them to a spectral chart, and record their results. Each group can report an approximate wavelength, and the average wavelengths for the class can be calculated for each color. One can then calculate the energies of each color, and plot them versus the observed color.

The size and construction of the forum is another key consideration. This concerns not only the size of the place you wish to perform the program or activity in, but also, the capacity, the presence or absence of bench ventilation, how much bench space is available, and the safety equipment and features of the area. This also refers to outside or inside activities. If one wishes to do a large scale emissions demo outside, one must consider the wind when choosing between using colored balloon explosions, which may be hard to ignite, or the chance that the methanol based flame demo is more dangerous in a cramped setting due to the accumulation of methanol fumes. Safety is a paramount concern. Some very colorful precipitation demos are beautiful, but both the products and the reactants have toxic properties, which may make them better suited as demos for younger crowds, but can work as activities for older students. Sometime one might decide to show video for a particularly dangerous or difficult to perform demo.

The setting under which a demo is to be performed can be casual, public, professional, as a formal lecture item, or as a teaching tool. What level of interest do you expect to generate? How much time is available for further investigation? Will you be seeing the audience again? Will you see them again as pupils or as observers? How much follow up are you able to do?

Theory

The use of emission spectra in the discussion of atomic properties, and excited states is well documented, and the emissions in this demonstration can be easily observed as well, both their similarities and their dissimilarities. The reason for those differences may be discussed, such as valence shell configuration, HOMO/LUMO energy gap, frequency and wavelength of the emitted light, and the idea of spectral analytical methods. How HOMO/LOMO (HOAO/LUAO) subject can be discussed with this demonstration. The material in the balloon, i.e. LiCl, is a metal salt in water/methanol. LiCl exists as a bonded Li^+ Cl^- pair, solvated by H_2O/methanol. When the balloon filled with hydrogen and the salt solution is ignited there are several major energy exchanges taking place. The first is during the conversion of H_2 (g) balloon + O_2 (g) atmosphere. The reaction is exothermic. The other two major energy exchanges are the heating of the gas during the reaction, the heating and vaporization of the water in the balloon, and the ionization of the electrons in the lithium salt.

The heating and vaporization of 1 g of water requires approximately 420 J*. The heating of the gaseous reactants is small due to their poor thermal conductivity in comparison to liquids present. The energy released during the combustion of hydrogen is sufficient to excite the electrons in the Li^+ ions. When the electrons in the Li^+ goes from ground state to its first excited state, it is actually moving to its first unoccupied frontier orbital, Li'^+. The electron cannot remain there for very long and soon returns to its ground state releasing its acquired energy in a photon of a specific wavelength corresponding to the energy difference between the ground and excited states.; the difference in the HOMO and LUMO since it is in the LiCl molecule, or at least in the Li^+

HOAO and LUAO, if the Li^+ is dissociated from its bonded state. Given the high ionization energy of alkali metal halides it may still be part of the ionic solid bonding at the time of excitation. Knowing the light is red yields an approximation for the wavelength for the emission, which can be verified via the ocean optics spectrometer at about 635 to 670 nanometers, and the energy of the photon corresponding to the $Li^+ \rightarrow Li'^+$ excitation state can be calculated using the following equation:

$$E= hc /\lambda$$

where h is Planck's constant (~ 6.634×10^{-34} J·S), c is the speed of light in a vacuum (~300,000,000 m/s) and lambda (λ) is the wavelength in meters.

The energy expressed can then be multiplied by Avogadro's number and converted to kJ to express the energy in kJ/mol. For other colors, i.e., sodium wavelength = ~587nm, Cu (II) = ~530nm, potassium = ~410 nm, etc. The colors and their approximate wavelengths appear in Table A.

Pedagogy:

In my presentation of the concept, I make a transparency of the visible spectrum with the assigned values for wavelengths* and then, either as a hands-on activity or as a demonstration the students observe the colored flames. I usually have the students perform this activity themselves followed by my repeating the flame tests as a demonstration. The following describes how this is done.

1) I set up each station as follows. Place five or six different colored flame test salts at four to six work stations. Each station should have thirty Q-tips®, a burner, six 50-ml Erlenmeyer flasks or cups with about 10-15 ml of the metal chloride salts in

aqueous solution in each container. Each station should have a printed visible spectrum with a good approximation of the colors versus the values given for wavelength.

2) I ask the students in groups to perform individual tests, identify the predominant color as compared to the chart for each substance tested, and record their best guess at the wavelength of that color. Each "test" consists of dipping the Q-tip® in the cup and holding the moistened end in the burner flame. An alcohol lamp can be used in place of the Bunsen burner but the colors are not as easily discernable and there are safety hazards associated with the use of alcohol lamps. The students keep a record of their observations.

You can either have the chemicals as unknowns, or clearly labeled. I like to give them the colored salts as unknowns and then repeat the experiment as a demonstration with known chemical identities. Some inferences can be made by noting which agent is responsible for the color change and how do you know that it is responsible. For example, igniting sodium chloride and noting the color followed by igniting sodium acetate and noting the color is the same for both compounds even though they have different anions can assist the students in seeing that the color of the flame is due to the presence of sodium in the flame. Of course the flame emissions of sodium chloride and potassium chlorides can further clarify for the students that the metals in the salts are responsible for the observed colors and not the chloride. If the chloride ions were responsible for the color then the two chloride salt solutions should emit the same color, and they clearly do not.

3) Then, in large group discussion, fair approximations of the wavelengths for each color emitted are determined based on the average recorded value from each group's

data. In this way I can either go on directly to the calculations, if students correctly demonstrated their understanding of the relationship between color and wavelength, or I can further discuss the idea of a reference standard or a calibration curve. I use the transparency of a graph of color versus the approximate wavelength. (When one does an experiment with an unknown concentration in the Beer's law setting, one develops a graph correlating known concentrations of the species of interest and then compares the absorbance reading of the unknown with those of the graph. This exercise is a corollary of that idea. The class gathers color versus wavelength data, and we get an average wavelength for each color. We can then talk about averages, and how one does an experiment multiple times to get better results.)

4) Once the average wavelengths are identified and the class is in agreement, the students can both calculate the corresponding energies and plot the corresponding energies versus the color, wavelength, or both. The corresponding frequencies for each color are then calculated followed by a discussion about the speed of light. In more advanced classes a discussion about the causes of the various colors can be held. However, even with elementary students one can present the idea of a ground state versus an excited state, and how fleeting the excited state can be.

*See Appendix A

Appendix A:

Some useful quantities for demonstration and discussion:

Planck's constant	$h = 6.634 \times 10^{-34}$ J·S
Speed of light	$c = \sim 3 \times 10^{8}$ m/s
Wavelength	λ (meters)
Energy of a photon of light of certain wavelength	$E = hc/\lambda$

- Relationship of speed of light to frequency and wavelength $c = \upsilon\lambda$ (υ is frequency, in sec $^{-1}$, and λ is the wavelength, in meters)
- Energy absorbed or emitted when a substance changes temperature $E = mc(t_f - t_i)$ (m is the mass of the substance, c is the specific heat capacity in joules per gram, and the term $(t_f - t_i)$ is the change in temperature, for non-phase change in Celsius, or degrees Kelvin)
- Energy change involved in changing between liquid and vapor states $E = m\Delta H_{vap}$ (ΔH_{vap} is the heat of vaporization. The value for water is 2256 J/g.)

Table A: Visible spectrum color versus wavelength [5]

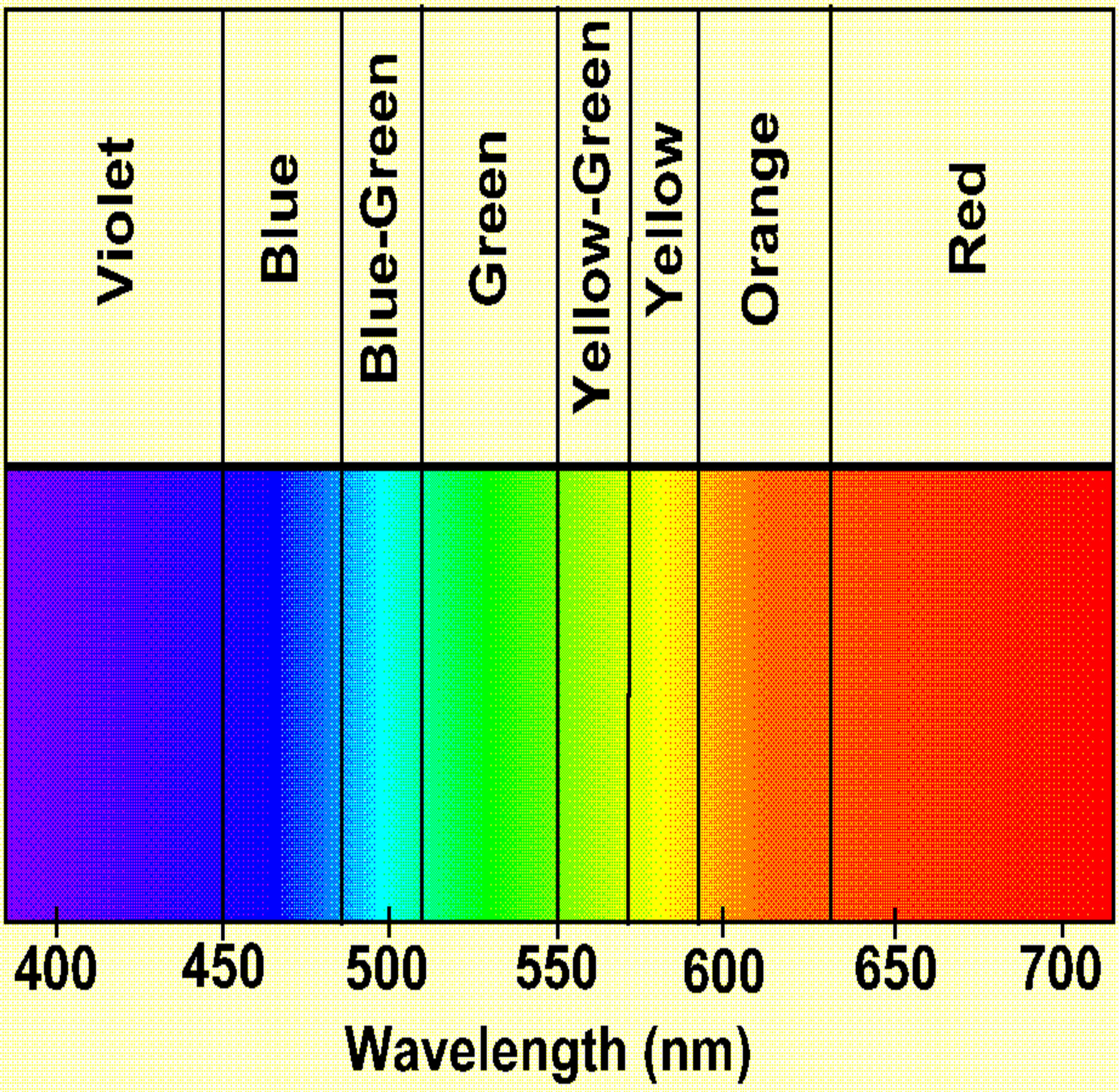

Figure 3: The visible spectrum

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5. Table A

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