The Ohio State University
Master Demonstration List
Undergraduate Chemistry

The Nature of Chemistry

1) Introduction to Chemistry – the following demos are useful as introductory illustrations to the world of chemistry
   a) **Apple Candle** – Light a "candle", let it burn a minute, blow it out, then eat it, wick and all. (The "candle" is an apple with an almond wick, but it looks exactly like a real candle.)
   b) **Burning Magnesium Ribbon** – Burn a piece of magnesium ribbon in air to produce MgO
   c) **Ethanol: Composition, Structure, and Reaction** – Display a flask of 95% ethanol and a ball-and-stick model of ethanol, then demonstrate burning ethanol in a watchglass to introduce the terms composition, structure, and reaction, which briefly summarize what the science of chemistry is all about.
   d) **H_2/O_2 Balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic (and loud) reaction that produces water.
   e) **Happy and Unhappy Balls** – Start with a pair of seemingly identical rubber balls to discuss potential and kinetic energy. One ball bounces like a superball, and the other absorbs energy and does not bounce.
   f) **Penny vs. Nitric Acid** – Read Ira Remson's own account of his first encounter with chemistry as you repeat his experiment by pouring concentrated nitric acid over a copper penny.
   g) **Rainbow Cups** – Add a colorless liquid to 6 "empty" beakers, producing the colors of the rainbow – use this demonstration to show how evaluation of observations and experimental results leads to hypotheses and further testing (the scientific method)
   h) **Ripping Pop Cans** – Ask a student to rip an aluminum pop can in half. When they cannot, demonstrate "proper" technique by effortlessly ripping a specially prepared can. Then hand them a prepared can and ask them to repeat the activity.

2) States of Matter
   a) **Halogens** – Display flasks containing the halogens chlorine, bromine, and iodine.
      i) On request, you can order a special flask of bromine that can be frozen in liquid nitrogen
   b) Use two sealed lucite boxes containing BBs on the overhead projector to illustrate some of the differences between gases, liquids, and solids
   c) **Changes of State** - Heat a beaker of ice on a hot plate during the lecture to show the changes from solid to liquid to gas
   d) **Changes of State** - Pour liquid nitrogen into a beaker to demonstrate a variety of physical changes: the boiling of N_2 (ℓ), the deposition of H_2O (g) as H_2O (s) on the outside of the beaker, and the melting of H_2O (s) to H_2O (ℓ) as the beaker eventually warms up again

3) Elements and Compounds
   a) **Elements and Compounds** – Show bottled samples of various elements and some compounds containing those elements (*e.g.* Mg, Cu, and S with $MgSO_4$ and $CuSO_4$)
   b) **Electrolysis of Water** – Electrolyze water (Na_2SO_4 solution with indicator) in the Hoffman apparatus to decompose it into its component elements, hydrogen and oxygen.
   c) **Reaction of Iron and Sulfur –** Lay a red-hot iron rod into a powdered mixture of the elements iron and sulfur to produce a non-magnetic compound, iron(**II**) sulfide, FeS (s)
      i) Pair with a heterogeneous mixture of iron and sulfur, and show how the mixture can be separated with a magnet

4) Mixtures
   a) **Heterogeneous and Homogeneous Mixtures** - Add solid NaCl to solid $CuSO_4$ in a beaker and stir – this is a heterogeneous mixture. Add water to the beaker and stir – a homogeneous mixture (a solution) results.

5) Physical and Chemical Changes
   a) **Burning Iron and Magnesium** – Contrast the results of holding an iron strip and a magnesium ribbon in a flame
   b) **Changes of state** – Heat a beaker of ice on a hot plate during the lecture to show the changes from solid to liquid to gas
   c) **Changes of State** - Pour liquid nitrogen into a beaker to demonstrate a variety of physical changes: the boiling of N_2 (ℓ), the deposition of H_2O (g) as H_2O (s) on the outside of the beaker, and the melting of H_2O (s) to H_2O (ℓ) as the beaker eventually warms up again
   d) **Combustion of Methane Bubbles** - Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students!
   e) **Dry Ice Sublimation** – Make the sublimation of dry ice "visible" by dropping a piece of dry ice in a beaker of water. (Optional: place a piece of dry ice in a glove, tie it off, and allow the dry ice to sublime, inflating the glove.)
   f) **Penny vs. Nitric Acid** – Read Ira Remson's own account of his first encounter with chemistry as you repeat his experiment by pouring concentrated nitric acid over a copper penny.
   g) **Smashing things with Liquid Nitrogen** – Demonstrate the coolant properties of liquid nitrogen by freezing a racquet ball or another object of your choice and smashing it.

6) Separation of Mixtures
   a) **Separation of a Mixture: Iron and Sulfur** – Separate iron filings and sulfur with a magnet

7) Units of Measurement
   a) Length – Show a meter stick
   b) Mass – Contrast 1 kilogram and 1 gram Te.
   c) Temperature
      i) Have a student read the temperature of ice water and boiling water with a digital thermometer in °C, then °F.

d) Volume
   i) Contrast the markings on a 1 L beaker and a 1 L graduated cylinder
   ii) Pour water from a 1 L graduated cylinder to a 1 dm cube, having first asked the class to predict the outcome, *e.g.* will the water overflow the cube.
   iii) Contrast 1 L, 100 mL and 10 mL graduated cylinders

e) Density
   i) Display three beakers containing 100 g of water, 100 g of sulfur, and 100 g of copper and lead the students to a qualitative understanding of density by asking questions about the three samples
   ii) **Density Demo – Candle** – Display a 1 L graduated cylinder with a colored candle stub floating at the half-way point in a clear colorless liquid, and ask the students to explain this phenomenon; then drop candle stubs into separate beakers of ethanol and water to show that the candle sinks in ethanol but floats in water – the cylinder contains water and ethanol in two layers, so the candle stays at the interface
   iii) Use **density decimeter cubes** of Mg, Al, Fe, and Pb to contrast the densities of these four metals
   iv) Pass several sets of **1" density cubes** of Al, Fe, Cu, and Pb around the class so students can qualitatively compare their densities
   v) **Density Demo – Silver** – Determine the mass of a "silver" cube on a balance which displays the mass on the overhead projector, then determine the volume of the cube by water displacement in a 250 mL graduated cylinder, and finally calculate the density to answer the question "Is this really a **silver** cube?" A table showing the density of several “silvery” metals is provided to help the class identify the cube as aluminum, not silver.
   vi) **Density of Coke vs. Diet Coke** – Drop unopened cans of regular Coke and Diet Coke into an aquarium filled with water. Coke sinks and Diet Coke floats – challenge the class to postulate an explanation.

8) Uncertainty in Measurement
   a) **Precision and Accuracy** – Using the document camera, measure a brass rod using rulers divided every 5 units, every 1 unit, and every 0.1 unit to introduce the concepts of precision and accuracy or significant figures

9) Dimensional Analysis
   a) Use (empty) egg cartons as a visual aid in a simple introduction to dimensional analysis using 12, a dozen, and/or a carton as the basis for unit factors
   b) Allow a few students to experience what 1 atm of pressure feels like by standing a special iron bar on their toes, then use dimensional analysis to relate the dimensions of the iron bar in inches and the density of iron in g/cm^3 to the pressure exerted (14.7 $pounds/in^2$)
   c) Contrast a U.S. soft drink can, which contains 12 fluid ounces (355 mL) with a European soft drink can, which contains 0.33 L, then use dimensional analysis to calculate the number of fluid ounces in the European soft drink can

Atoms, Molecules, and Ions

1) The Atomic Theory of Matter
    a) **Conservation of Mass** – Demonstrate the law of conservation of mass with the colorful reaction of $Co(NO_3)_2$ (aq) and Na_2CO_3 (aq), using the document camera to project the digital readout of a balance on the screen before and after the reaction.

2) The Discovery of Atomic Structure
    a) **Cathode Ray Tube and Pinwheel** – Demonstrate the deflection of an electron beam with a magnet (CRT), and the particle nature of electrons by using a beam of electrons to spin a pinwheel.
        i) Optional - Use the Geiger counter to show students the radiation that the CRT emits
    b) **Detection of Radioactivity** – Use a Geiger counter to demonstrate the radioactivity (or lack thereof) of several substances, including NaI, NaC and uranium salts. A sheet of lead is provided to display the ability of lead to block radiation.
    c) **Gas Discharge Tubes of the Noble Gases** – show that different gases give different colors when subjected to an electric discharge (H_2 tube also available)

3) The Periodic Table
    a) Display samples of various elements from different parts of the periodic table
        i) We can choose a variety of elements for you, or you may specify
    b) Pass samples of C (as charcoal), Si, Sn, and Pb around the class so students can compare and contrast a nonmetallic element, a metalloid, and two metallic elements from Group 4A of the periodic table (**H1/2**)
    c) Pass vials containing C, S, Si, Sb, Mg, and Co around the class so students can compare and contrast two nonmetallic elements, two metalloids, and two metallic elements from across the periodic table (**H1/2**)
    d) **Halogens** – Display flasks containing the halogens chlorine, bromine, and iodine.
        i) On request, you can order a special flask of bromine that can be frozen in liquid nitrogen
    e) **Periodic Properties** – Add pieces of Li, Na, K, Mg, and Ca, to beakers of water to observe the reactivity of metals from different parts of the periodic table.

4) Molecules and Molecular Compounds
    a) Show ball-and-stick models of simple molecules such as H_2, H_2O, H_2O_2, O_2, O_3, CO_2, CO, CH_4, C_2H_2, C_6H_6, and C_2H_5OH; you may also wish to show models of some elements such as Cl_2, P_4, and S_8. (Please indicate desired models upon ordering.)

5) Ions and Ionic Compounds
    a) Show large open lattice models such as NaCl, CsCl, ZnS, $CaCO_3$, CaF_2 and/or a space-filling model of NaCl. (Please indicate which models you want when you make your request. Extended lattices are available for NaCl and CsCl if you prefer.)

6) Naming Inorganic Compounds
   a) Use samples of an element and some of its compounds as a visual aid when discussing nomenclature. For example:
      i) Large cube of Fe (s) or a jar of Fe strips, clear jars of iron(II) sulfate and iron(III) sulfate
      ii) Jars of Cu (s), S (s), and $CuSO_4$ (s) with or without a stoppered tube of O_2 (g)
   b) Show reagent bottles or clear jars of salts containing "ous" or "ic" ions, polyatomic ions, or hydrates as a visual aid when discussing nomenclature of these compounds. For example:
      i) $CuCl_2$, CuCl, and CuI
      ii) $FeSO_4$ and $Fe_2(SO_4)_3$

Chemical Formulas and Equations

1) Chemical Equations – choose an attention-getting reaction to introduce writing and balancing equations:
   a) **Burning Magnesium Ribbon** – Burn a piece of magnesium ribbon in air to produce MgO
   b) **Combustion of Methane Bubbles** – Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students! If a more intense demo is needed, try Methane Bubbles XTREME.
   c) **Potassium and Water** – Drop a piece of potassium into an aquarium containing water and phenolphthalein to produce H_2 (g) and KOH (aq) – the heat of reaction ignites the H_2 (g) and a lavender flame is observed (from the K^+), while the indicator turns pink from the formation of KOH.

2) Combination Reactions
   a) **Reaction of Iron and Sulfur** – Lay a red-hot iron rod into a powdered mixture of the elements iron and sulfur to produce a non-magnetic compound, iron(II) sulfide,FeS (s).
   b) **Reaction of Zinc and Sulfur** – Lay a red-hot iron rod into a powdered mixture of zinc and sulfur to produce a shower of sparks and ZnS.
   c) **Combustion of Sulfur in Oxygen** – Burn sulfur in air enriched with O_2 to produce SO_2(g), then dissolve the product in water containing universal indicator to show that SO_2 is an acidic oxide – this shows how acid rain results from burning high sulfur coal.
   d) **H_2/O_2 Balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic (and loud) reaction to produce water.
   e) **Burning Magnesium Ribbon** – Burn a piece of magnesium ribbon in air to produce MgO

3) Decomposition Reactions
   a) **Electrolysis of Water** – Electrolyze water (dilute Na_2SO_4 solution with indicator) in the Hoffman apparatus to decompose it into its component elements, hydrogen and oxygen.
   b) **Ammonium Dichromate Volcano** – Ignite a pile of $(NH_4)_2Cr_2O_7$ volcano in an aquarium to produce N_2 (g), H_2O (g), and Cr_2O_3 (s). The resulting reaction looks very much like an erupting volcano.

4) Combustion in Air
   a) **Burning Magnesium Ribbon** – Burn a piece of magnesium ribbon in air to produce MgO
   b) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the explosive flammability of organic vapors.
   c) **Ether Fire** – Demonstrate the combustion of ether by allowing the vapor to flow from a can down an inclined trough to a candle, resulting in a vapor flashback fire.
   d) **Combustion vs. Fire** – Contrast the burning of ethanol with the burning of magnesium ribbon. One is combustion and the other is not. Does fire always mean combustion?

5) Avogadro's Number and the Mole
   a) **Avogadro's Number and the Mole** - Illustrate the constant mass ratio of equal numbers of atoms of different elements by counting out equal numbers of large and small balls and comparing their masses in "ball units" analogous to the amu.
   b) Use empty egg cartons to develop an analogy between 12 and a dozen as compared to Avogadro's number and the mole.
   c) **Mole Samples** – Show mole samples of various elements and compounds in bottles
   d) **Avogadro's Law** – Three flasks containing equal amounts of acetic acid are fitted with balloons containing different amounts of $NaHCO_3$; mix the reagents by lifting and shaking the balloons: the balloons will inflate with CO_2 to a volume proportional to the number of moles produced, in accordance with Avogadro's law.

6) Empirical Formulas from Analyses
   a) Show ball-and-stick models to match empirical and molecular formulas derived in lecture:
      i) C_6H_6 and/or C_2H_2 from CH
      ii) Any cycloalkane or alkene from CH_2

7) Quantitative Information from Balanced Equations – Choose an attention-getting reaction to lead into stoichiometry problems based on that reaction
   a) **Combustion of Methane Bubbles** – Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students! If a more intense demo is needed, try Methane Bubbles XTREME.
   b) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the explosive flammability of organic vapors.
   c) **Potassium and Water** – Drop a piece of potassium into an aquarium containing water and phenolphthalein to produce H_2 (g) and KOH (aq) – the heat of reaction ignites the H_2 (g) and a lavender flame is observed (from the K^+), while the indicator turns pink from the formation of KOH.

Quantities of Reactants and Products

1) **Avogadro's Law** – Three flasks containing equal amounts of acetic acid are fitted with balloons containing different amounts of $NaHCO_3$. Mix the reagents by lifting and shaking the balloons, starting the reaction. The balloons will inflate with CO_2 to a volume proportional to the number of moles produced, in accordance with Avogadro's law.
2) **H_2/O_2 Balloon Series** – Light a series of balloons containing different ratios of H_2 (g) and O_2 (g): pure H_2, 2:1, 1:1, 1:2, pure O_2. The loudest bang occurs when the ratio is stoichiometric, 2:1.
3) **Acid-Base Titration** – Add 1 M NaOH (aq) to a solution of 0.1 M HCl and phenolphthalein to show the endpoint.

Energy and Chemical Reactions

1) Enthalpies of Reaction – Endothermic Reactions
   a) **An Endothermic Reaction** – Shake solid $Ba(OH)_2 \cdot 8\ H_2O$ with solid NH_4NO_3 to produce an aqueous mixture of $Ba(NO_3)_2$ (s) and NH_3 (aq). The reaction is endothermic enough to freeze the flask to a wet piece of cardboard. Alternatively, a digital thermometer can be used to record the temperature change.
   b) **Hot and Cold Packs** – Combine NH_4NO_3 (s) and water in a Ziploc bag to make an instant "cold pack".
   c) **Thionine Reaction** – Hold a solution of thionine and $FeSO_4$ in front of a bright light to show the reduction of thionine from a violet form to a colorless form; this is an endothermic reaction that absorbs light energy.

2) Enthalpies of Reaction – Exothermic Reactions (**Note:** A digital thermometer can be supplied to measure ΔT in some reactions. Other reactions may require a hood.)
   a) **Combustion of Methane Bubbles** – Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students! If a more intense demo is needed, try Methane Bubbles XTREME.
   b) **Methane Bubbles XTREME** – Ignite a tower of methane-filled soap bubbles to produce a pillar of flame 3-5m high. As seen on Mythbusters!
   c) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the explosive flammability of organic vapors
   d) **Crystallization of Supersaturated Sodium Acetate Solution** – Add a tiny crystal of sodium acetate to a 2 L flask of a super-saturated solution to cause the solid to crystallize out, leaving almost no liquid – this demonstration is beautiful and dramatic, as well as being quite exothermic.
   e) **Hot and Cold Packs** – Combine $CaCl_2$*(s)* and water to make an instant "hot pack"
   f) **Luminol** – Pour solutions of luminol and H_2O_2 into a tall glass spiral to produce a beautiful chemiluminescent reaction. The light-emitting species is the dicarboxylate ion, aminophthalate, the product of the oxidation of luminol with H_2O_2
      i) Alternative is Luminol Ammonia Fountain with luminol instead of water and phenolphthalein
   g) **Sound** – Tap a very small pile of red phosphorus and $KClO_3$ with a hammer to show a reaction that produces light, sound, and heat, and recreates on a larger scale the reaction that occurs when you strike a match.
   h) **Glycerin and Potassium Permanganate** – Pour glycerin over $KMnO_4$ (s) to initiate a spontaneous combustion accompanied by smoke, sparks, and a lavender flame.
      i) **Note:** While this reaction is spontaneous, the rate depends on surface area, and it may take up to 2 minutes to react.
   i) **Burning Magnesium Ribbon** – Burn a piece of magnesium ribbon in air to produce MgO.

j) **H_2/O_2 Balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic reaction to produce water.

k) **Thermite** – Perform the thermite reaction, in which Al and Fe_2O_3 react to produce molten iron.

   i) **NOTE: 48 hour notice is required for this demonstration**

l) **Combustion of Candy** – Contrast the oxidation of sucrose in the body (by eating some candy) with the oxidation of sucrose by $KClO_3$ (as shown by dropping some candy into molten $KClO_3$, producing steam and a lavender flame. Body temperature is ~37°C, and the melting point of $KClO_3$ is 368°C.

3) Bond Enthalpies – Demonstrate one of the interesting reactions listed below, then draw Lewis structures of selected reactants and/or products

a) **Graham's Law of Diffusion** – Allow concentrated NH_3 and concentrated HCl to vaporize and meet in a horizontal glass tube, forming a ring of NH_4Cl.

   HCl (g) + NH_3 (g) → NH_4Cl (s)

b) **Combustion of Methane Bubbles** – Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students! If a more intense demo is needed, try Methane Bubbles XTREME.

   CH_4 (g) + 2 O_2 (g) → CO_2 (g) + 2 H_2O (g)

c) **Combustion of Candy** – Contrast the oxidation of sucrose in the body (by eating some candy) with the oxidation of sucrose by $KClO_3$ (as shown by dropping some candy into molten $KClO_3$, producing steam and a lavender flame. Body temperature is ~37°C, and the melting point of $KClO_3$ is 368°C.

   $C_{12}H_{22}O_{11}$ (s) + 8 $KClO_3$ (ℓ) □ 8 KCl (s) + 12 CO_2 (g) + 11 H_2O (g)

d) **Ammonia Fountain** – Show the solubility of NH_3 (g) in H_2O due to hydrogen-bonding

   NH_3 (g) + H_2O (ℓ) □ NH_4^+ (aq) + OH^- (aq)

e) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the explosive flammability of organic vapors

   C_2H_5OH(g) + 3 O_2(g) □ 2 CO_2(g) + 3 H_2O(g)

f) **H_2/O_2 balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic reaction to produce water.

   2 H_2(g) + O_2(g) □ 2 H_2O(g)

Electronic Structure of Atoms

1) Line Spectra and the Bohr Model
   a) Use a prism on the overhead projector to disperse white light into a continuous spectrum.
   b) **Gas Discharge Tubes of the Noble Gases** – show that different gases give different colors when subjected to an electric discharge (H_2 tube also available)
   c) **Atomic Spectra: Flame Tests** – Introduce various metal salts (e.g. NaCl, SrCl2, CuCl2) into an open flame to produce brilliant colors associated with exciting metal cations.
   d) **Fireworks Balloons** – Show the brilliant colors of metal cations in a fun way by igniting hydrogen balloons containing metal salts.
   e) **Pickle Electrocutor** – Run an electric current through a dill pickle using a special apparatus to show students that the brine in the pickle conducts electricity, and the current heats and excites the sodium atoms present to a higher energy level. The emission of photons as the sodium returns to the ground state accounts for the yellow color of the light.
   f) **Diffraction of Light** – Use a laser pointer and a slide containing eight different arrays of dots to simulate x-ray diffraction experiments.

2) Representations of Orbitals
   a) Show models of *s*, *p*, *d*, and *f* orbitals
   b) Show large styrofoam balls in two sizes to represent the 1*s* and 2*s* orbitals and a large styrofoam model of the 2*p* subshell
   c) Show a large styrofoam model of the 2*s* and 2*p* orbitals nestled together in contrasting colors
   d) Show a styrofoam hemisphere painted to represent a cross-sectional view of the probability distribution of electron density inside a 2*s* orbital

3) Electron Spin and the Pauli Exclusion Principle
   a) **Paramagnetic and Diamagnetic Salts** – Provide experimental evidence of electron spin by bringing a powerful magnet close to suspended test tubes of $MnSO_4$, $FeSO_4$, $NiSO_4$, and $ZnSO_4$ to show the different responses due to different numbers of unpaired electrons

4) Molecular Orbital Theory
   a) To help students visualize how p orbitals interact to form molecular orbitals, use a large styrofoam model of the p subshell and a large styrofoam model of a p orbital (show two p orbitals end-to-end and then two parallel p orbitals along the y axis or the z axis), or use two small models (side-by-side) of the p subshell (px, py, and pz)
   b) Contrast models of NH_2—NH_2, NH=NH, and N≡N to show the decreasing N–N bond length as the bond order increases
   c) **Paramagnetic O_2** – Demonstrate the paramagnetism of liquid oxygen by pouring first $N_2(\ell)$, then $O_2(\ell)$ between the poles of a powerful magnet displayed by the document camera.

The Periodic Table

1) Metals, Nonmetals, and Metalloids
   a) Pass vials containing C, S, Si, Sb, Mg, and Co (**Metals, Nonmetals, Metalloids**) around the class so students can compare and contrast two nonmetallic elements, two metalloids, and two metallic elements from across the periodic table
   b) **Acidic and Basic Oxides** – Dissolve several oxides (CaO, ZnO, CO_2, P_4O_{10}) in water containing universal indicator to show a range of basic and acidic oxides

2) Group Trends for the Active Metals
   a) **Periodic Properties** – Add pieces of Li, Na, K, Mg, and Ca, to beakers of water and phenolphthalein to observe the reactivity of metals from different parts of the periodic table. If desired, you can add HCl to those beakers where no reaction occurred
   b) **Potassium and Water** – Drop a piece of potassium into an aquarium containing water and phenolphthalein to produce H_2 (g) and KOH (aq) – the heat of reaction ignites the H_2 (g) and a lavender flame is observed (from the K+), while the indicator turns pink from the formation of KOH.

3) Group Trends for selected Nonmetals
   a) **H_2/O_2 balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic reaction to produce water.
   b) Group 6A: The Oxygen Group.
      i) **Elephant Toothpaste** – Demonstrate the decomposition of 30% H_2O_2 in the presence of dishwashing liquid and KI, producing an upsurge of steaming foam.
      ii) **Combustion of Sulfur in Oxygen** – Burn sulfur in air enriched with O_2 to produce SO_2(g), then dissolve the product in water containing universal indicator to show that SO_2 is an acidic oxide – this shows how acid rain results from burning high sulfur coal
   c) Group 7A: The Halogens.
      i) **Halogens –** Display flasks containing the halogens chlorine, bromine, and iodine
   d) Group 8A: The Noble Gases.
      i) **Gas Discharge Tubes of the Noble Gases** – show that different gases give different colors when subjected to an electric discharge (H_2 tube also available)

Covalent Bonding

1) Illustrate covalent bonds using large styrofoam models that represent overlap of two s orbitals and overlap of two p orbitals

2) Show ball-and-stick models of simple molecules or polyatomic ions; you may wish to show a ball-and-spring model to remind students that covalent molecules are not rigid

3) Show space-filling models of molecules such as H_2O, CO_2, or CH_4

4) Drawing Lewis Structures
   a) Demonstrate one of the interesting reactions listed below (or another reaction of your choice), then draw Lewis structures of selected reactants and/or products; ball-and-stick models of reactants and products can also be provided upon request
      i) **Combustion of Methane Bubbles** – Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students! If a more intense demo is needed, try Methane Bubbles XTREME.
      $CH_4\ (g) + 2\ O_2\ (g) \rightarrow CO_2\ (g) + 2\ H_2O\ (g)$
      ii) **Ammonia Fountain** – Show the solubility of NH_3 (g) in H_2O due to hydrogen-bonding.
      $NH_3\ (g) + H_2O\ (\ell) \rightarrow NH_4^+\ (aq) + OH^-\ (aq)$

   b) Show ball-and-stick models of selected molecules or ions, with short sticks for lone pairs
      i) Select from NH_3, ClO_2^-, $COCl_2$, CO, H_2CO, NO_3^-

5) Resonance Structures – Show ball-and-stick models of selected molecules that display resonance
   a) Select from benzene, CO_3^{2-}, NO_3^-, SO_3, NO_2^-, SO_2

6) Strengths of Covalent Bonds – Perform one of the following demonstrations and use bond energies to estimate the ΔH of reaction
   a) **H_2/O_2 balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic reaction to produce water.
   b) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the flammability of organic vapors.

7) Single, Double, and Triple Covalent Bonds
   a) Show models of Cl_2, O_2, and N_2 to contrast single, double, and triple bonds
   b) Contrast models of NH_2—NH_2, NH=NH, and N≡N to show the decreasing N–N bond length as the bond order increases
   c) Show orbital overlap models of ethane, ethene, and ethyne (acetylene)

Molecular Geometry and Structure

1) The VSEPR model
   a) Use ball-and-stick models to illustrate the VSEPR shapes and sub-shapes. Please specify if you want sub-shapes.
      i) Linear
      ii) Trigonal Planar – trigonal planar and bent 120°
      iii) Tetrahedral – tetrahedral, trigonal pyramidal, and bent 109°
      iv) Trigonal Bipyramidal – trigonal bipyramidal, see-saw, T-shaped
      v) Octahedral – octahedral, square pyramidal, square planar
   b) Use 5 sets of balloons to illustrate the VSEPR shapes: linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral.
   c) Start with a set of 4 balloons tied together and pop them one by one to show the geometries that arise naturally from repulsion of "electron pairs."
   d) Display a lucite tetrahedron containing a model of methane to illustrate the terms "tetrahedron" and "tetrahedral"

2) Molecular Shape and Molecular Polarity
   a) **Polarity and Geometry** – Show the dependence of dipole-dipole forces on geometry by contrasting the effect of a charged rod on streams of H_2O and "CCl_4" (actually hexane) flowing from burets.
   b) Contrast ball-and-stick models of different compounds to relate the geometry of these molecules to their dipole moments
      i) CCl_4, CH_3Cl, and CH_2Cl_2
      ii) NH_3, SO_3, and H_2CO
      iii) H_2O and CO_2

3) Covalent Bonding, Orbital Overlap, Sigma and Pi Bonds
   a) Illustrate covalent bonds using large styrofoam models that represent overlap of two s orbitals and overlap of two p orbitals
   b) Bring pairs of orbitals close together (such as s and s, s and p, two p orbitals end-to-end, and finally two parallel p orbitals) to show how σ and π bonds are formed

4) Hybrid Orbitals
   a) Use individual models of s, p, and d orbitals and models of sp, sp^2, sp^3, "dsp^3"*, and "d^2sp^3"*hybrid orbitals in appropriate combinations to show which atomic orbitals produce the various hybrid orbitals
      *not real

5) Multiple Bonds
   a) Bring pairs of orbitals close together (such as s and s, s and p, two p orbitals end-to-end, and finally two parallel p orbitals) to show how ʃ and π bonds are formed
   b) Use a large styrofoam model of the p subshell and a large styrofoam model of a p orbital to bring pairs of p orbitals close together (two p orbitals end-to-end and then two parallel p orbitals along the y axis or the z axis) to show how p orbitals form ʃ and π bonds
   c) Show ball-and-stick models of simple molecules which have double and triple bonds, showing only the sigma bonds initially; then add pairs of semicircular rings to represent the pi bonds
      i) C_2H_4 and C_2H_2
      ii) CO and CO_2
   d) Contrast orbital overlap models and ball-and-stick models of C_2H_6, C_2H_4, and C_2H_2 to illustrate the hybrid orbitals and atomic orbitals used to form the various sigma and pi bonds;
   e) Contrast models of 1,2-dichloroethane, cis-1,2-dichloroethylene, and trans-1,2-dichloroethylene
   f) Delocalized π bonding
      i) Use a styrofoam model of benzene to show delocalization in the π bonding orbital
      ii) Show ball-and-stick models to discuss delocalized π bonding
         (1) Choose from benzene, CO_3^{2-}, NO_3^-, SO_3, NO_2^-, and SO_2

6) Second-Row Diatomic Models
   a) To help students visualize how p orbitals interact to form molecular orbitals, use a large styrofoam model of the p subshell and a large styrofoam model of a p orbital (show two p orbitals end-to-end and then two parallel p orbitals along the y axis or the z axis), or use two small models (side-by-side) of the p subshell (px, py, and pz)
   b) Contrast models of NH_2—NH_2, NH=NH, and N≡N to show the decreasing N–N bond length as the bond order increases
   c) **Paramagnetic O_2** – Demonstrate the paramagnetism of liquid oxygen by pouring first $N_2(\ell)$, then $O_2(\ell)$ between the poles of a powerful magnet on the overhead projector or document camera.

Gases

1) Characteristics of Gases
    a) Use two sealed Lucite boxes containing BBs on the overhead projector to illustrate some of the properties that distinguish gases from liquids and solids

2) Pressure
    a) **Squashing Pop Cans with Atmospheric Pressure** – use atmospheric pressure to squash pop cans
    b) Allow a few students to experience what 1 atm (14.7 psi) "feels" like by resting an iron bar one inch square and 54 inches long on their toes

3) The Pressure-Volume Relationship: Boyle's Law
    a) **Marshmallow Snowman** – Demonstrate the effect a decrease in *P* has on *V* by placing a marshmallow snowman in a bell jar and then evacuate the jar.
    b) **Potato Rifle** – Use a "potato rifle" (long pipe with two potato cores) to dramatically demonstrate the effect a decrease in volume has on pressure.
    c) **Boiling Water at Room Temperature** – Show water boiling at room temperature in a beaker in an evacuated bell jar, then put your hand in the water after boiling to convince students of its low temperature. (Upon request, we can give you a clean beaker and tap water if you want to drink the boiled water)

4) The Temperature-Volume Relationship: Charles' Law
    a) **Charles' Law** – Pour liquid nitrogen over a balloon to show that a decrease in T is accompanied by a decrease in V.

5) The Quantity-Volume Relationship: Avogadro's Law
    a) **Avogadro's Law** – Three flasks containing equal amounts of acetic acid are fitted with balloons containing different amounts of $NaHCO_3$; mix the reagents by lifting and shaking the balloons: the balloons will inflate with CO_2 to a volume proportional to the number of moles produced, in accordance with Avogadro's law
    b) Display the 22.4 L box to illustrate molar volume at STP
    c) **Molar Volume** – Place 28 g of liquid nitrogen in a special cube balloon and tie off; after the N_2 has vaporized, compare the volume of the bag to the 22.4 L box

6) Kinetic-Molecular Theory
    a) Use the Molecular Motion demonstrator to show the random motion of particles and the increase in velocity with increasing temperature.
    b) Use the "gas" lucite box of BBs to show the random motion of gas molecules.
    c) Use magnetized steel balls and the Molecular Motion Demonstrator on the overhead to show how attractive forces cause a gas to liquefy at lower temperatures.

7) Molecular Effusion and Diffusion
   a) **Vanilla Balloons** – Pass around a few balloons containing a potent osmophore (vanilla) and ask students to identify the odor; the odor is detectable because of the diffusion of vanillin molecules through pores in the balloon.
   b) **Graham's Law** – Allow concentrated NH_3 and concentrated HCl to vaporize and meet in a horizontal glass tube, forming a ring of NH_4Cl.

8) Miscellaneous Gas Demonstration – Fluidity of Gases
   a) **Lake Nyos Demo** – Pour CO_2 (g) down an enclosed set of steps to extinguish candles on each step, demonstrating the fluidity of gases, and recreating (on a small scale) a tragic natural disaster

Liquids, Solids, and Materials

1) A Molecular Comparison of Gases, Liquids and Solids
   a) Use two sealed Lucite boxes containing BBs on the overhead projector to illustrate some of the properties that distinguish gases from liquids and solids
   b) Use magnetized steel balls in the molecular motion demonstrator on the overhead projector first to show the random motion of particles in a gas and then to show how attractive forces cause a gas to liquefy at lower temperatures
   c) **Halogens** – Display flasks containing the halogens chlorine, bromine, and iodine.
      i) On request, you can order a special flask of bromine that can be frozen in liquid nitrogen

2) Intermolecular Forces – Use ball-and-stick models of compounds used in the following demonstrations to relate geometry and structure to intermolecular forces:
   a) **Polarity and Geometry** – Show the dependence of dipole-dipole forces on geometry by contrasting the effect of a charged rod on streams of H2O and "CCl4" (actually hexane) flowing from burets.
   b) **Polarity and Solubility** – Add acetone to a saturated solution of $CuSO_4(aq)$ causing $CuSO_4(s)$ to crystallize out – the solubility of $CuSO_4$ decreases as the polarity of the solvent is decreased
   c) Use pairs of space-filling models of n-pentane and isopentane to show that increased branching increases compactness, decreases polarizability, and decreases London forces
      i) Pair with velcro models to simulate the strength of IMF between branched and un-branched hydrocarbons
   d) **Ammonia Fountain** – Show the solubility of NH_3 (g) in H_2O due to hydrogen-bonding. Alternative: Luminol Ammonia Fountain
   e) Show **PAIRS** of ball-and-stick models of HF, H_2O, and NH_3, with plain rods for lone pairs to help illustrate hydrogen bonding. Optional: single models of H_2S, PH_3, and SiH_4 to show molecules that do not exhibit hydrogen bonding

3) The Liquid State
   a) Viscosity
      i) **Viscosities of Liquids** – Compare the viscosity of several liquids via the viscosity apparatus - hexane, ethanol, water, and glycerin.
   b) Surface Tension
      i) **Surface Tension and Micelle Formation** – Demonstrate the disruption of the surface tension of water by detergent by sprinkling baby powder in a large dish of water and then touching the water lightly with a wood stick dipped in detergent. Use the overhead projector or document camera to display the settling of powder through the water.
      ii) **Gooyuk/Ooblak** – Pass around bowls of a cornstarch-water mixture; the properties of this non-Newtonian fluid challenge our traditional definitions of liquid and solid (and is awesome).

4) Phase Changes
   a) **Changes of state** – heat a beaker of ice on a hot plate during the lecture to show the changes from solid to liquid to gas.
   b) **Halogens** – Display flasks containing the halogens chlorine, bromine, and iodine.
      i) On request, you can order a special flask of bromine that can be frozen in liquid nitrogen
   c) **Dry Ice Sublimation** – Make the sublimation of dry ice "visible" by dropping a piece of dry ice in a beaker of water. (Optional: place a piece of dry ice in a glove, tie it off, and allow the dry ice to sublime, inflating the glove.)
   d) **Changes of state** – Pour liquid nitrogen into a beaker to demonstrate a variety of phase changes: the boiling of $N_2(\ell)$, the deposition of $H_2O(g)$ as $H_2O(s)$ on the outside of the beaker, and the melting of $H_2O(s)$ to $H_2O(\ell)$ as the beaker eventually warms up again

5) Vapor Pressure
   a) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the flammability of organic vapors.
   b) **Ether Fire** – Demonstrate the high vapor pressure of diethyl ether by allowing the vapor to flow from a can down a trough to a candle, resulting in a vapor flashback fire.
   c) **Comparing the Vapor Pressure of Two Liquids** – Show the effect of intermolecular forces on the vapor pressure of liquids by contrasting the vapor pressures of two isomers, diethyl ether and 1-butanol.
   d) **Vapor Pressure: Evaporation** – Show the effect of intermolecular forces on vapor pressure by letting students make streaks of water, methanol, and acetone on the blackboard and then observing the relative rates of evaporation.

6) Vapor Pressure and Boiling Point
   a) **Boiling Water at Room Temperature** – Show water boiling at room temperature in a beaker in an evacuated bell jar, then put your hand in the water after boiling to convince students of its low temperature. (Upon request, we can give you a clean beaker and tap water if you want to drink the boiled water)

7) Phase Diagrams
   a) **Effect of Pressure on the Melting Point of Ice** – Hang a wire weighted at both ends over a cylinder of ice; eventually the wire passes through the ice and the weights fall, leaving the ice intact. The ice below the wire melts due to pressure from the weights, and the water above the wire refreezes as the pressure is relieved.
      i) **NOTE: This demo requires 48 hour notice**
   b) **Dry Ice Sublimation** – Make the sublimation of dry ice "visible" by dropping a piece of dry ice in a beaker of water. (Optional: place a piece of dry ice in a glove, tie it off, and allow the dry ice to sublime, inflating the glove.)
   c) **Triple Point of CO_2** – Demonstrate the existence of three phases of CO_2 at the triple point by adding crushed dry ice to a clear acrylic tube fitted with a pressure gauge and a release valve

8) **Liquid Crystals** – Demonstrate the color changes initiated by temperature changes on two Mylar sheets covered with black ink and liquid crystals; placing a hand on the film or "writing" on it with a damp Kimwipe are two ways to bring about color changes. Compare the temperatures needed to effect change between different sheets

9) Structure of Solids
   a) Crystalline and Amorphous Solids – contrast a piece of charcoal, a large quartz crystal, a piece of pumice, and a polished quartz crystal (SiO_2).
   b) Unit Cells and Crystal Lattices
      i) Display ball-and-stick models of simple cubic, BCC, and FCC unit cells (Optional: Display space-filling models and/or extended lattice models).
      ii) Stack several space-filling FCC cubes or BCC cubes to show how unit cells form an extended lattice.
      iii) Pass around small unit cell models of simple cubic, BCC, and FCC.
      iv) DIsplay unit cell models of NaCl or CsCl.
      v) Illustrate hexagonal and cubic close-packed structures with layers of ping pong balls.
      vi) **Diffraction of Light** – Use a laser pointer and a slide containing eight different arrays of dots to simulate x-ray diffraction experiments.

10) Metallic Solids
   a) Show models of Cu, Mg, and Fe lattices.

11) Ionic Solids
   a) Display models of NaCl, CsCl, ZnS, CaF_2, and TiO_2
   b) Display a model of $CaCO_3$

12) Molecular Solids
   a) Show models of CO_2 (s) and H_2O (s) lattices.
   b) Show pairs of Darling models of benzene and toluene to illustrate the effect of structure and symmetry of molecules on their melting and boiling points.

13) Covalent Network Solids – show large models of different arrangements of pure carbon: graphite, diamond, and/or C_{60} (buckminsterfullerene).

14) Polymeric Solids
   a) **Disappearing Styrofoam Cup** – Make a styrofoam (expanded polystyrene) cup disappear by placing it in a dish of acetone.
   b) **Nylon 6-10** – Demonstrate the polymerization of hexamethylenediamine with sebacoyl chloride to produce the polyamide Nylon 6-10.
   c) Cross-linking Polymers
      i) **Disposable Diaper Demo** – See how much water you can add to a super-absorbent disposable diaper, then cut open another diaper to show the super-absorbent powder, Water Lock J-550, which is polysodium acrylate cross-linked with starch; the original polymer results from multiple addition reactions of the alkene functional groups of acrylic acid molecules. (One diaper holds 1 L of water!)
      ii) **Slime!** – make a cross-linked gel by mixing solutions of polyvinyl alcohol and borax; use this demo to relate concepts such as polymers and hydrogen-bonding to a commercial product students are familiar with.
   d) Ceramics
      i) **The Meissner Effect** – Demonstrate the levitation of a tiny magnet over a cooled superconductor

Chemical Kinetics

1) The Effect of Concentration on Rate
   a) **Iodine Clock Reaction** – Perform the iodine clock reaction with three different initial concentrations of IO_3-
   b) **Combustion of Methane Bubbles** and **H_2/O_2 balloon** – Contrast the rate of combustion of methane-filled soap bubbles with the rate of combustion of a 1:2 mixture of methane and oxygen (actually H_2 and O_2) in a balloon; note that activation energy in the form of a flame is required to initiate both reactions.
   c) **Steel Wool Reaction Rates** – Contrast the burning of steel wool in air and in a flask charged with O_2 (g)

2) The Effect of Surface Area on Rate
   a) **Lycopodium** – Contrast the rate of combustion of a pile of lycopodium powder versus the exploding paint can
      i) Alternative: Blow lycopodium powder into a candle flame
   b) **Surface Area Reaction Rates – Iron Strip and Powder** – Contrast the results of holding a strip of iron in a burner versus squirting powdered iron into the burner flame
   c) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the flammability of organic vapors.
      i) Contrast this with the burning of ethanol in a watch glass

3) Temperature and Rate
   a) **Alka Seltzer at Three Temperatures** –Three students add Alka-seltzer tablets to flasks containing water at different temperatures and quickly seal the flasks with stoppers fitted with balloons, which will inflate at different rates
   b) **Light Sticks** – Immerse light sticks in hot and cold water to show variation in rates depending on temperature
   c) **Combustion of Candy** – Contrast the rate of oxidation of sucrose in the body (by eating some candy) with the oxidation of sucrose by $KClO_3$ (as shown by dropping some candy into molten $KClO_3$, producing steam and a lavender flame. Body temperature is ~37C, and the melting point of $KClO_3$ is 368C

4) The Collision Model and Activation Energy
   a) Use ball-and-stick models to show how orientation of molecules influences the effectiveness of a collision
      i) H_2, I_2, and 2 HI
   b) **H_2/O_2 balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic reaction to produce water.

5) Reaction Mechanisms
   a) **Briggs-Rauscher Oscillating Reaction** – Introduce the mystery of mechanisms with the Briggs-Rauscher Oscillating Reaction.
   b) **A Simple Oscillating Reaction** – Based on the B-Z oscillating reaction, solutes are added in sequence to produce a red-colorless oscillating reaction. This is an alternative to the Briggs-Rauscher reaction, as it does not produce iodine.
   c) **Oscillating Methanol Explosion** – A platinum wire is heated in a flask containing methanol, creating a small explosion that deprives the local environment of oxygen. As more oxygen diffuses in, the platinum wire heats again, causing another explosion
   d) **Reaction Intermediates** – Add $FeCl_3$ (aq) to $Na_2S_2O_3$ (aq) on the overhead projector, giving rise to the black intermediate $FeS_2O_3^+$, which forms and disappears, leaving colloidal sulfur as the final product.
   e) To show the unlikelihood of a termolecular collision, give colored foam balls to three students and challenge them to throw the balls so that all three collide simultaneously

6) Catalysis
   a) **Catalysis of a Reaction** – Demonstrate the catalysis of the H_2O_2 decomposition of NaK-tartrate with Co^2+. Adding Co^2+ turns the solution pink, but the solution quickly turns dark green as it begins to react vigorously. At the end of the reaction, the pink color is restored showing regeneration of the catalyst;
   b) **Genie in a Bottle** – Use MnO_2 to catalyze the decomposition of 30% H_2O_2, producing a large cloud of hot water vapor. The heat generated is intense enough to shrink the 2 L bottle used for the demo
   c) **Elephant Toothpaste** – Demonstrate the decomposition of 30% H_2O_2 in the presence of dishwashing liquid and KI, producing an upsurge of steaming foam.
   d) **Catalytic Decomposition of H_2O_2** – Compare and contrast several different catalysts, including chicken liver, used for the decomposition of hydrogen peroxide
      i) **NOTE**: this demo requires 48 hours notice

Chemical Equilibria

1) **Dynamic Equilibrium** – Introduce the concept of dynamic equilibrium by having two students bail beads between two clear boxes, one empty, one full, until equilibrium is achieved. This demo helps students view dynamic equilibrium as a state where the **rates** of forward and reverse reactions are equal, not where the **amounts** of reactants and products are equal.

2) Le Chatelier's Principle: Changes in Reactant or Product Concentration
   a) **Le Chatelier's Principle: Iron(III) Thiocyanate Equilibria** – Apply stress to the $Fe^{3+} + SCN^- \rightleftharpoons FeSCN^{2+}$ system in five different ways to show the equilibrium shifts accompanying changes in the concentration of reactants
   b) **Chromate-Dichromate Equilibrium** – Show the pH dependence of the $CrO_4^{2-}/Cr_2O_7^{2-}$ system
   c) **Cobalt Complexes and Temperature v2.0** – Demonstrate effects of concentration and temperature changes on the $Co(H_2O)_6^{2+}/CoCl_4^{2-}$ equilibrium

3) Le Chatelier's Principle: Changes in Temperature
   a) **Effect of Temperature on $NO_2 \leftrightarrow N_2O_4$ Equilibrium** – Immerse sealed tubes of NO_2/N_2O_4 in hot and cold water to show how temperature shifts the equilibrium position and to show the reversibility of the shift; red-brown NO_2 predominates at high temperatures and colorless N_2O_4 at lower temperatures
   b) **Cobalt Complexes and Temperature v2.0** – Demonstrate effects of concentration and temperature changes on the $Co(H_2O)_6^{2+}/CoCl_4^{2-}$ equilibrium

Solutes and Solutions

1) General Properties of Aqueous Solutions
   a) **Conductivity Tester Demos 1**
      i) **Sugar and Salt** – Use two conductivity testers with light bulbs to contrast the conductivity of d-H_2O, sugar solution, and NaCl (aq).
      ii) **Strong and Weak Acids and Bases** – Use two conductivity testers with light bulbs to contrast the conductivity of weak and strong electrolytes: acetic acid and HCl (aq), and/or NH_3 (aq) and NaOH (aq)
   b) Display a wave bottle, which contains oil and colored water to illustrate a mixture that does not form a solution.

2) Precipitation Reactions – Perform one (or more) of these reactions and then explain the results using solubility rules, or tell the students the proposed reactants, have them predict the results using solubility rules, then perform the demo to test their hypotheses.
   a) **Precipitation Reactions** – mix two pairs of solutions to show that some combinations produce precipitate and some give no net reaction.
      i) $CoCl_2$ and Na_2CO_3 vs. $CoCl_2$ and Na_2SO_4
      ii) KI and $Pb(NO_3)_2$ vs. KI and $NaNO_3$
   b) **Lead Nitrate and Potassium Iodide - Solid and Liquid Reactions** – Mix equal amounts of $Pb(NO_3)_2$ (s) and KI (s) in a test tube and shake thoroughly to show formation of yellow PbI_2 (s). Contrast this with the results of mixing $Pb(NO_3)_2$ (aq) and KI (aq)

3) Molarity
   a) Use a 1 L bottle of 1.0 M $CuSO_4$ (or another colored solution) as a visual aid when defining molarity, then pour 100 or 200 mL into a 1 L beaker and calculate the number of moles of solute in the beaker

4) Dilution
   a) Add water to a measured volume of 1.0 M $CuSO_4$ (or another colored solution) in a beaker, and calculate the concentration of the dilute solution

5) The Solution Process
   a) Add a few crystals of $KMnO_4$ to water on the overhead projector so students can observe the process of dissolution
   b) **Hot and Cold Packs** – Demonstrate dramatic differences in heats of solution by dissolving NH_4NO_3(s) in water in a Ziploc bag to make an instant "cold pack" and $CaCl_2$(s) in water to make an instant "hot pack", then pass the bags around the class

6) Saturated Solutions and Solubility
    a) **Crystallization of Supersaturated Sodium Acetate Solution** – Add a tiny crystal of sodium acetate to a 2 L flask of a super-saturated solution to cause the solid to crystallize out, leaving almost no liquid – this demonstration is beautiful and dramatic, as well as being quite exothermic. Alternatively you can pour the solution slowly over a single crystal to build up a column of solid sodium acetate
    b) **Polarity and Solubility** – Add acetone to a saturated solution of $CuSO_4(aq)$ causing $CuSO_4(s)$ to crystallize out – the solubility of $CuSO_4$ decreases as the polarity of the solvent is decreased

7) Factors Affecting Solubility: Solute-Solvent Interactions
    a) **Ammonia Fountain** – Show the solubility of NH_3 (g) in H_2O due to hydrogen-bonding. Alternative: Luminol Ammonia Fountain
    b) **Negative Volume of Mixing** – Mix ethanol and colored water in a graduated cylinder to demonstrate their miscibility and negative volume of mixing due to hydrogen bonding.
    c) **Miscible vs. Immiscible Liquids** – Mix ethanol and colored water in one beaker and hexane and colored water in another to demonstrate miscibility and immiscibility due to differences in the intermolecular forces of alcohols as the size of the alkyl group increases. The demo can be repeated with butanol and water, to show that, as carbon chain length increases, polarity of alcohols decreases.
    d) **Polarity and Solubility** – Add acetone to a saturated solution of $CuSO_4(aq)$ causing $CuSO_4(s)$ to crystallize out – the solubility of $CuSO_4$ decreases as the polarity of the solvent is decreased
    e) **Like Dissolves Like** – Contrast the solubility of $I_2(s)$ and $CuCl_2(s)$ in both water and hexane in large test tubes

8) Factors Affecting Solubility: Temperature Effects
    a) **Solubility and Temperature** – Heat two flasks, one containing saturated $Ca(C_2H_3O_2)_2$ and the other saturated KNO_3; the calcium acetate will crystallize out and the potassium nitrate dissolves
    b) **Glittering Shower of Lead Iodide Crystals** – Cool a hot colorless flask containing a solution of $PbI_2(aq)$ to produce a glittering shower of yellow $PbI_2(s)$ crystals.

9) Colligative Properties: Lowering Vapor Pressure
    a) **Osmotic Pressure** – small dialysis bags containing equimolar solutions of isopropanol and $CaCl_2$ are attached to long glass tubes; immerse the bags in distilled water to illustrate osmosis and to show that osmotic pressure depends on the number of particles in solution.
    b) **Vapor Pressure Lowering of Solutions** – Demonstrate the lower vapor pressure of solutions by knocking over a medicine cup of salt in a sealed manometer containing water.

10) Colloids
   a) **The Tyndall Effect** – Demonstrate the Tyndall effect and simulate a sunset on the overhead projector by reacting $Na_2S_2O_3$ with HCl to produce a colloidal suspension of sulfur.

Acids and Bases

1) Strong and Weak Acids and Bases
   a) **Acids and Bases** – Add drops of 0.1 M NaOH to a magnetically stirred solution of water and universal indicator until it turns blue, then add drops of 0.1 M HCl until the solution turns orange.
   b) **Conductivity Tester Demos (1 and 2)**
      i) **Strong and Weak Acids and Bases** – Contrast the extent of ionization in weak and strong acids and bases using the lightbulb conductivity apparatus.
      ii) **HCl and NaOH** – Use the conductivity tester to show that HCl and NaOH conduct but H_2O does not. Add Yamada Universal Indicator to show that HCl is an acid, NaOH is a base, and H_2O is neutral. Finally mix equal amounts of HCl and NaOH to produce a neutral solution – will it still conduct electricity?
      iii) **Glacial Acetic Acid** – Use the conductivity tester to show that glacial (concentrated) acetic acid does not conduct electricity. Add some d-H_2O, and the solution will now conduct electricity. Why?

2) Neutralization Reactions and Salts
   a) **Acid-Base Titration** – Add NaOH (aq) to a solution of HCl and phenolphthalein
   b) **Electrolytic Titration of $Ba(OH)_2$ with H_2SO_4** – Add H_2SO_4 gradually to $Ba(OH)_2$ (aq) in which the conductivity tester is immersed; the light bulb gradually dims as the endpoint is reached, because the products are H_2O and insoluble $BaSO_4$
   c) **Conductivity Testers 2**
      i) **Acetic Acid and Ammonia** – use the conductivity tester to show that acetic acid and ammonia are weak electrolytes. Have students predict the results of mixing equal amounts of ammonia and acetic acid then test the conductivity of the mixture – does it still conduct electricity? Does the bulb glow dimly or brightly? Why?
   d) **Milk of Magnesia Demo** – Add HCl gradually to Milk of Magnesia in a beaker (an aqueous suspension of $Mg(OH)_2$) to show the dissolution of $Mg(OH)_2$ as the pH decreases

3) Acid-Base Reactions with Gas Formation
   a) **Gas Forming Reaction Demos**
      i) $CaCO_3$ and HCl – Add H_2O to solid $CaCO_3$ to show that it is insoluble, then add HCl(aq): CO_2(g) is generated and the solid dissolves completely if enough HCl is added
      ii) Na_2CO_3 vs. $NaHCO_3$ (sodium carbonate vs. sodium bicarbonate) – Attach separate balloons containing Na_2CO_3 and $NaHCO_3$ to flasks of acetic acid (or HCl). Lift the balloons to begin the acid-base reactions and compare the amounts of gas generated.
      iii) Acetic Acid vs. HCl – Attach balloons containing separate portions of $NaHCO_3$ to flasks containing HCl and acetic acid. Lift the balloon to begin the acid-base reactions and collect the gas generated, and compare the size of the balloons to talk about strong and weak acids.
      iv) Buffer vs. Acid – Attach balloons containing separate portions of $CaCO_3$ to flasks containing acetic acid and acetic acid/sodium acetate buffer. Lift the balloon to begin the acid-base reactions and collect the gas generated, and compare the size of the balloons to talk about the relative rates of reaction.
   b) **Alka Seltzer Poppers** – React an Alka Seltzer tablet with water in a film canister (placed inside a tube with a ping pong ball on top) to show that the production of gas is so great it pops the top off the film canister and launches the ball
   c) **Acid Rain** – Add a small piece of blackboard chalk ($CaCO_3$) to 6 M HCl (aq) to show a more vigorous acid-base reaction.

4) Brønsted-Lowry Acids and Bases
   a) Toss a racquet ball labeled H^+ to a student, then hold up your hand so the student throws the ball back; now ask "Do you know what you've done? . . .You've made an acid of yourself!" This should indelibly imprint the Brønsted -Lowry concept of an acid as a proton donor in your students' minds
   b) Show ball-and-stick models of Brønsted -Lowry acid and base reactants and products, such as the following:
      i) NH_3 and H_2O to give NH_4^+ and OH^-
      ii) H_2O and H_2O, to give H_3O^+ and OH^-

5) Acidic, Basic, and Amphoteric Oxides
   a) **Acidic and Basic Oxides** – Dissolve several oxides (CaO, ZnO, CO_2, P_4O_{10}) in water containing universal indicator to show a range of basic and acidic oxides
   b) **ZnO and HCl –** Add ZnO to water to show that it does not dissolve appreciably, then add acid to show that a basic oxide reacts with (and therefore dissolves in) an acid
   c) **Amphoteric Hydroxides: Al^{3+} and Fe^{3+}** – add NaOH to samples of $Al(NO_3)_3$ and $Fe(NO_3)_3$ to form insoluble metal hydroxides, then add HNO_3 and more NaOH to different samples of each to identify which metal hydroxides are amphoteric

6) The pH Scale
   a) **Yamada and Dry Ice** – Add a chunk of dry ice to a 2 L cylinder containing a basic solution and Yamada universal indicator; the dry ice gradually acidifies the solution causing the color to change in the order purple, blue, green, yellow, orange
   b) **Yamada and Breath** – Have a student wearing a white or light-colored shirt use a straw to blow into water containing Yamada universal indicator or to observe the color change accompanying the reaction

7) Strong Acids and Bases
   a) **Conductivity Tester Demos 1**
      i) **Strong and Weak Acids and Bases** – Use two conductivity testers with light bulbs to contrast the conductivity of weak and strong electrolytes: acetic acid and HCl (aq), and/or NH_3 (aq) and NaOH (aq)
   b) **Reactions of Strong and Weak Acids** – Contrast the rates of reaction of 1 M HCl and 1 M $HC_2H_3O_2$ with $NaHCO_3$

8) Acid-Base Properties of Salt Solutions
   a) **Acidity and Basicity of Salts** (Formerly Hydrolysis of Salts) – Dissolve various salts in water containing Yamada indicator to demonstrate their acidity or alkalinity in solution.
      i) Usual compounds given: NH_4NO_3, NaF, NaCl, Na_2CO_3, $Fe(NO_3)_3$, $NaHSO_4$
      ii) Salts available: NaF, NaCl, Na_2CO_3, $NaHSO_4$, $NaNO_2$, NH_4NO_3, $(NH_4)_2CO_3$, NH_4CH_3COO, $AlCl_3$
      iii) Metals available: $Fe(NO_3)_3$, $Al_2(SO_4)_3$, $Pb(NO_3)_2$, $Zn(NO_3)_2$, $Ca(NO_3)_2$, $NaNO_3$, KNO_3, $AgNO_3$, $Cu(NO_3)_2$

9) Lewis Acids and Bases
   a) Toss a racquet ball, now with a lone pair, to a student, then hold up your hand so the student throws the ball back. Whoever tosses this ball makes an acid of the person who catches it, as that person becomes an electron pair acceptor.
   b) **Graham's Law** – Allow concentrated NH_3 and concentrated HCl to vaporize and meet in a horizontal glass tube, forming a ring of NH_4Cl
   c) Show ball-and-stick models of Lewis acid and base reactants and products:
      i) $NH_3 + BF_3 \rightarrow F_3B{:}NH_3$
      ii) $AlCl_3 + Cl^- \rightarrow AlCl_4^-$
      iii) $NH_3 + HCl \rightarrow NH_4^+$ and Cl^-
   d) **Acidity and Basicity of Salts** (Formerly Hydrolysis of Salts) – Dissolve various nitrates in water containing Yamada indicator to demonstrate their acidity or alkalinity in solution.
      i) Usual compounds given: $Pb(NO_3)_2$, $Zn(NO_3)_2$, $AgNO_3$, $Cu(NO_3)_2$, $Fe(NO_3)$
      ii) Nitrates available: NH_4NO_3, $NaNO_3$, $Pb(NO_3)_2$, $Zn(NO_3)_2$, $Ca(NO_3)_2$, $Fe(NO_3)_3$, KNO_3, $AgNO_3$, $Cu(NO_3)_2$

Oxidation-Reduction Reactions

1) Oxidation Numbers – demonstrate one of the interesting reactions listed below (or another reaction of your choice), then calculate the oxidation number of selected atoms and/or balance the equation by the change-in-oxidation-number method
   a) **Ammonium Dichromate Volcano** – Ignite an $(NH_4)_2Cr_2O_7$ volcano in an aquarium to produce N_2(g), H_2O(g), and Cr_2O_3(s)
   b) **Combustion of Magnesium in Dry Ice** – Demonstrate the combustion of magnesium in dry ice, resulting in a luminous effect as well as the production of MgO and elemental Carbon (requires 48 hours notice).
   c) **Reaction of Zinc and Sulfur** – Lay a red-hot iron rod into a powdered mixture of zinc and sulfur to produce a shower of sparks and ZnS
   d) **Multiple Oxidation States of Vanadium** – Shake a solution of ammonium meta-vanadate with a Zn-Hg amalgam to reduce the vanadium from +5 to +4 to +3 to +2 with different colors at each stage
   e) **Oxidation of Alcohols** – Demonstrate the oxidation of ethanol with $K_2Cr_2O_7$ on the overhead projector; the alcohol solution changes from orange to green to blue as the Cr(VI) is reduced; this reaction is the basis for the Breathalyzer test

2) Oxidation of Metals by Acids and Salts – choose one of these metal A/metal ion B (or H^+) reactions to demonstrate a displacement reaction; you may also wish to show metal B immersed in metal ion A to show that the reverse reaction does not occur
   a) **Ripping Pop Cans** – Ask a student to rip an aluminum pop can in half. When they cannot, demonstrate "proper" technique by effortlessly ripping a specially prepared can. Then hand them a prepared can and ask them to repeat the activity.
   b) **Metal Redox Reactions** – Compare redox reactions between metals and metal ions to see which is spontaneous
      i) **Copper and Aluminum** – immerse a strip of Cu in $Al(NO_3)_3$ (aq) and compare to a ball of aluminum foil in $CuCl_2$ (aq) (the second is also the reaction used to prepare the Ripping Pop Cans demo)
      ii) **Copper and Iron** – Immerse a large iron nail in $CuSO_4$ (aq) and compare to a strip of Cu in $FeSO_4$(aq).
      iii) **Copper and Silver** – Immerse a strip of copper in $AgNO_3$ (aq) and compare to “silver” wire in $CuSO_4$ (aq)
      iv) **Copper and Zinc** – Immerse a strip of Cu in $ZnSO_4$ (aq) and compare to a strip of Zn in $CuSO_4$ (aq).
      v) **Zinc and Lead** – Immerse a strip of Zn in $Pb(NO_3)_2$ (aq), and compare to a strip of Pb in $ZnSO_4$ (aq).
      vi) **Copper Star Oxidation** – Immerse copper wire in $AgNO_3$ (aq). This is best done on the document camera over a lengthy period of time so students can observe the continuing reaction.

c) **Zinc and HCl** – Immerse mossy zinc in HCl (aq) and use a lighted splint to ignite the hydrogen gas produced.
d) **Aluminum and NaOH** – Immerse aluminum foil in NaOH (aq) and use a lighted splint to ignite the hydrogen gas produced.

Aqueous Equilibria

1) Buffer Action and Buffer Capacity
   a) **Buffer Demo** – Contrast the buffer capacity of water, 1 M $CH_3COOH/NaCH_3COO$, and 0.1 M $CH_3COOH/NaCH_3COO$ by adding increments of 6 M HCl to each in the presence of an indicator
   b) **Gas Forming Reactions**
      i) Buffer vs. Acid – Attach balloons containing separate portions of $CaCO_3$ to flasks containing acetic acid and acetic acid/sodium acetate buffer. Lift the balloon to begin the acid-base reactions and collect the gas generated, and compare the size of the balloons to talk about the relative rates of reaction.

2) Acid-Base Titrations
   a) **Acid-Base Titration** – Add 1 M NaOH (aq) to a solution of 0.1 M HCl and phenolphthalein to show the endpoint.
   b) **Titration of a Triprotic Acid**

3) The Common Ion Effect
   a) **Common Ion Effect Demos**
      i) **HCl/NaCl** – add concentrated HCl (aq) to saturated NaCl(aq) to cause precipitation of NaCl (s)
      ii) **Acetic Acid/Sodium Acetate** - Add sodium acetate to a solution of acetic acid and indicator; the color change accompanying the change in pH shows the equilibrium shift caused by the common ion effect
4) Dissolving Precipitates
   a) **Common Ion Effect Demos**
      i) **Solubility of Magnesium Hydroxide –** Mix 1 M $MgCl_2$ (aq) and 1 M NH_3 (aq) to form $Mg(OH)_2$ (s). Add NH_4Cl (s) and NH_4+ removes OH- to reform NH_3, so $Mg(OH)_2$ dissolves to replace OH-
   b) **Milk of Magnesia Demo** – Add HCl gradually to Milk of Magnesia in a beaker (an aqueous suspension of $Mg(OH)_2$) to show the dissolution of $Mg(OH)_2$ as the pH decreases
   c) **Amphoteric Hydroxides: Al^{3+} and Fe^{3+}** – add NaOH to samples of $Al(NO_3)_3$ and $Fe(NO_3)_3$ to form insoluble metal hydroxides, then add HNO_3 and more NaOH to different samples of each to identify which metal hydroxides are amphoteric
5) Complex Ion Formation
   a) **Complex Ion Formation Ni or Cu** – Show color change associated with formation of complex ions. In a tall graduated cylinder of Cu^{2+} or Ni^{2+} aqueous solutions, carefully add 6 M NH_3 to create a layering effect of $[M(H_2O)_6^{2+}$ / $M(OH)_2$ / $[M(NH_3)_4^{2+}]$
      i) $CuSO_4$ gives a blue / white / dark blue layering
      ii) $Ni(NO_3)_2$ gives a green / white / dark blue layering
   b) **Precipitates and Complexes of Silver** – add a series of 7-9 reagents to Ag+(aq) in succession to produce a series of precipitates and complexes.

Thermodynamics

1) Heat Capacity and Specific Heat
   a) **Money to Burn** – Soak a dollar bill in a water-alcohol mixture and then light it with a match; the high specific heat of water keeps the combustion temperature low enough to prevent burning the bill

2) Calorimetry
   a) Display a coffee cup calorimeter and a bomb from a bomb calorimeter when you discuss constant-pressure calorimetry versus constant-volume calorimetry
   b) **Coffee Cup Calorimeter** – Use a coffee cup calorimeter, a thermometer, and measured amounts of $CaCl_2$ and water to demonstrate how a calorimeter works

3) Spontaneous Processes
   a) Display a large box lid containing white balls on one side and colored balls on the other side, then walk around the room carrying the box to demonstrate the spontaneous mixing of the balls. Is the "reaction" spontaneously reversible? Compare to separate beakers of colored balls.

4) Spontaneity, Enthalpy, and Entropy – even endothermic reactions can be spontaneous if they have a large positive ΔS
   a) Set a beaker of ice in front of the class to be observed as the lecture progresses
   b) **An Endothermic Reaction** – Shake solid $Ba(OH)_2 \cdot 8\ H_2O$ with solid NH_4NO_3 to produce an aqueous mixture of $Ba(NO_3)_2$ (s) and NH_3 (aq). The reaction is endothermic enough to freeze the flask to a wet piece of cardboard. Alternatively, a digital thermometer can be used to record the temperature change

Electrochemistry

1) Oxidation States and Oxidation-Reduction Reactions
   a) **Reaction of Iron and Sulfur** – Lay a red-hot iron rod into a powdered mixture of the elements iron and sulfur to produce a non-magnetic compound, iron(II) sulfide, FeS (s).
   b) **Reaction of Zinc and Sulfur** – Lay a red-hot iron rod into a powdered mixture of zinc and sulfur to produce a shower of sparks and ZnS
   c) **Burning Magnesium Ribbon** – Burn a piece of magnesium ribbon in air to produce MgO
   d) **Metal Redox Reactions** – Compare redox reactions between metals and metal ions to see which is spontaneous
      i) **Copper and Aluminum** – immerse a strip of Cu in $Al(NO_3)_3$ (aq) and compare to a ball of aluminum foil in $CuCl_2$ (aq) (the second is also the reaction used to prepare the Ripping Pop Cans demo)
      ii) **Copper and Iron** – Immerse a large iron nail in $CuSO_4$ (aq) and compare to a strip of Cu in $FeSO_4$(aq).
      iii) **Copper and Silver** – Immerse a strip of copper in $AgNO_3$ (aq) and compare to “silver” wire in $CuSO_4$ (aq)
      iv) **Copper and Zinc** – Immerse a strip of Cu in $ZnSO_4$ (aq) and compare to a strip of Zn in $CuSO_4$ (aq).
      v) **Zinc and Lead** – Immerse a strip of Zn in $Pb(NO_3)_2$ (aq), and compare to a strip of Pb in $ZnSO_4$ (aq).
      vi) **Copper Star Oxidation** – Immerse copper wire in $AgNO_3$ (aq). This is best done on the document camera over a lengthy period of time so students can observe the continuing reaction.
   e) **Combustion of Magnesium in Dry Ice** – Demonstrate the combustion of magnesium in dry ice, resulting in a luminous effect as well as the production of MgO and elemental Carbon (requires 48 hours notice)
   f) **Zinc and HCl** – Immerse mossy zinc in HCl (aq) to produce $ZnCl_2$ (aq) and H_2 (g). Use a lighted splint to ignite the hydrogen gas produced.

2) Balancing Oxidation-Reduction Equations – use one of these demonstrations to show that H+ or OH^- is necessary to balance some redox reactions; when the reactants in the first and second demos are mixed, the expected product will not appear until acid is added
   a) **Balancing a Redox Equation: MnO_4^-/NO_2^-** – Reduce pink MnO_4- with NO_2- in aqueous solution to produce colorless Mn^{2+}
   b) **Aluminum and NaOH** – Add aluminum foil to NaOH (aq) to produce $Al(OH)_3$ (aq) and H_2 (g) Use a lighted splint to ignite the hydrogen gas produced.

3) Voltaic Cells
   a) **Zinc and Iodine Reaction** – Reaction of $Zn + I_2 \rightarrow ZnI_2$ – Show the reaction of Zn (s) with I_2 (s) in water, which is exothermic enough to produce a puff of the violet vapor of I_2.
      i) Optional: If you want to spend the time, the product can be electrolyzed to produce a Zn-plated cathode and I_3^- at the anode and use it to do work.
   b) **Copper/Zinc Voltaic Cell** – Demonstrate a copper/zinc voltaic cell turning a motor to show that a spontaneous reaction can be harnessed to do work.

4) The Nernst Equation
   a) **Concentration Cell** – Set up a concentration cell with 1 M Cu^{2+} on the bottom and 0.01 M Cu^{2+} on the top with copper plates immersed in the solutions as electrodes; the voltage read from a multimeter should be close to 59 mV as predicted by the Nernst equation

5) Batteries and Fuel Cells
   a) Pass around a disassembled 9 V battery to show that it is comprised of six 1.5 V cells

6) Electrolysis
   a) **Electrolysis of Water** – Electrolyze water (dilute Na_2SO_4 solution with indicator) in the Hoffman apparatus to decompose it into its component elements, hydrogen and oxygen. If desired, you can test the H_2 (g) and/or O_2 (g) produced with a flame and a glowing splint, respectively.

Nuclear Chemistry

1) **Detection of Radioactivity** – Use a Geiger counter to demonstrate the radioactivity (or lack thereof) of several substances, including NaI, NaC and uranium salts. A sheet of lead is provided to display the ability of lead to block radiation.

2) Nuclear Fission
   a) Display one liter (1 kg) of water and a cube of aluminum the size of 1 kg of uranium (3.75 cm on a side), both to contrast the densities of water and uranium and to accompany a calculation of the amount of energy released from fission of 1 kg of uranium

Main Group Elements

1) Periodic Trends
   a) Display samples of various elements from different parts of the periodic table (We can choose a variety of elements for you, or you may specify)
   b) Pass samples of C (as charcoal), Si, Sn, and Pb around the class so students can compare and contrast a nonmetallic element, a metalloid, and two metallic elements from Group 4A of the periodic table (**O1**)
   c) Pass vials containing C, S, Si, Sb, Mg, and Co around the class so students can compare and contrast two nonmetallic elements, two metalloids, and two metallic elements from across the periodic table (**O1**)

2) Hydrogen
   a) Properties of Hydrogen.
      i) **H_2/O_2 Balloon** – Ignite a balloon filled with a stoichiometric mixture of hydrogen and oxygen to show the extremely exothermic reaction to produce water.
      ii) Use a discharge tube of H_2 to show the color emitted
   b) Production of Hydrogen:
      i) **Zinc and HCl** – Immerse mossy zinc in HCl (aq) to produce $ZnCl_2$ (aq) and H_2 (g). Use a lighted splint to ignite the hydrogen gas produced.
      ii) **Potassium and Water** – Drop a piece of potassium into an aquarium containing water and phenolphthalein to produce H_2 (g) and KOH (aq) – the heat of reaction ignites the H_2 (g) and a lavender flame is observed (from the K+), while the indicator turns pink from the formation of KOH.
      iii) **Aluminum and NaOH** – Add aluminum foil to NaOH (aq) to produce $Al(OH)_3$ (aq) and H_2 (g) Use a lighted splint to ignite the hydrogen gas produced.

3) Group 1 and Group 2: The Alkali and Alkaline Earth Metals
   a) **Periodic Properties** – Add pieces of Li, Na, K, Mg, and Ca, to beakers of water to observe the reactivity of metals from different parts of the periodic table. If desired, you can add HCl to those beakers where no reaction occurred.
   b) **Potassium and Water** – Drop a piece of potassium into an aquarium containing water and phenolphthalein to produce H_2 (g) and KOH (aq) – the heat of reaction ignites the H_2 (g) and a lavender flame is observed (from the K+), while the indicator turns pink from the formation of KOH.
   c) **Combustion of Magnesium in Dry Ice** – Demonstrate the combustion of magnesium in dry ice, resulting in a luminous effect as well as the production of MgO and elemental Carbon (requires 48 hours notice).

4) Group 13/III: The Boron Family
   a) **Slime!** – make a cross-linked gel by mixing solutions of polyvinyl alcohol and borax; use this demo to relate concepts such as polymers and hydrogen-bonding to a commercial product students are familiar with.

5) Group 14/IV: The Carbon Family
   a) Elemental Forms of Carbon.
      i) Show models of graphite, diamond, and C_{60}, buckminsterfullerene
      ii) Pass around samples of powdered graphite and charcoal
   b) Oxides of Carbon.
      i) **Lake Nyos Demo** – Pour CO_2 (g) down an enclosed set of steps to extinguish candles on each step, demonstrating the fluidity of gases, and recreating (on a small scale) a tragic natural disaster
      ii) **Production and Combustion of CO** – Add concentrated H_2SO_4 to formic acid to produce H_2O and CO(g), then ignite the CO(g) to show that it burns with a blue flame.
         (1) Optional: you can also show that, at extremely low pH, phenolphthalein turns orange
   c) Carbonic Acid and Carbonates.
      i) Add dry ice, $NaHCO_3$, and Na_2CO_3 to beakers of universal indicator to determine the acidic or basic character of each compound
      ii) **Acid Rain** – Add a small piece of blackboard chalk ($CaCO_3$) to 6 M HCl (aq) to see the dissolution of a carbonate in acid.
   d) Other Group 14/IV Elements
      i) Pass samples of Si, Sn, and Pb around the class.
      ii) Contrast a chunk of the metalloid Si with a wafer of refined silicon used in electronic devices.
      iii) Pass small vials containing samples of Si, Ge, Sn, and Pb around the class so students can compare and contrast two metalloids and two metallic elements from the same group.
      iv) Display a large polished crystal of SiO_2 and pumice (another form of SiO_2).
      v) Display the lucite tetrahedron and/or a set of four tied balloons to illustrate the silicate tetrahedron as a building block of silicate minerals.

6) Group 15/V: The Nitrogen Family
   a) Nitrogen
      i) **Smashing things with Liquid Nitrogen** – Demonstrate the coolant properties of liquid nitrogen by freezing a racquet ball or another object of your choice and smashing it.
      ii) **Effect of Temperature on $NO_2 \leftrightarrow N_2O_4$ Equilibrium** – Immerse sealed tubes of NO_2/N_2O_4 in hot and cold water to show how temperature shifts the equilibrium position and to show the reversibility of the shift; red-brown NO_2 predominates at high temperatures and colorless N_2O_4 at lower temperatures
      iii) **Penny vs. Nitric Acid** – Read Ira Remson's own account of his first encounter with chemistry as you repeat his experiment by pouring concentrated nitric acid over a copper penny.
   b) Other Group 15/V Elements
      i) Pass around sealed vials containing small samples of P, As, Sb, and Bi.
      ii) **Sound** – Tap a very small pile of red phosphorus and $KClO_3$ with a hammer to show a reaction that produces light, sound, and heat, and recreates on a larger scale the reaction that occurs when you strike a match.
      iii) **Acidic and Basic Oxides** – Dissolve P_4O_{10} in water containing universal indicator to show an acidic solution, H_3PO_4 is produced

7) Group 16/VI: The Oxygen Family
   a) Production of Oxygen.
      i) **Electrolysis of Water** – Electrolyze water (dilute Na_2SO_4 solution with indicator) in the Hoffman apparatus to decompose it into its component elements, hydrogen and oxygen. If desired, you can test the H_2 (g) and/or O_2 (g) produced with a flame and a glowing splint, respectively.
   b) Oxides.
      i) **Acidic and Basic Oxides** – Dissolve several oxides (CaO, ZnO, CO_2, P_4O_{10}) in water containing universal indicator to show a range of basic and acidic oxides
      ii) **ZnO and HCl –** Add ZnO to water to show that it does not dissolve appreciably, then add acid to show that a basic oxide reacts with (and therefore dissolves in) an acid
   c) Peroxides and Superoxides.
      i) **Genie in a Bottle** – Use MnO_2 to catalyze the decomposition of 30% H_2O_2, producing a large cloud of hot water vapor. The heat generated is intense enough to shrink the 2 L bottle used for the demo.
   d) Other Group 16/VI Elements
      i) Show a model of S_8 and/or a jar of powdered sulfur
      ii) **Combustion of Sulfur in Oxygen** – Burn S_8 in air enriched with O_2 to produce SO_2 (g), then dissolve the SO_2 in H_2O containing an indicator to show that an acid is produced ($SO_2 + H_2O \rightarrow H_2SO_3$); this demonstration shows how acid rain results from burning high sulfur coal
      iii) **Black Carbon Snake** – Show the dehydrating properties of concentrated H_2SO_4 by

adding it to sugar in a tall beaker to produce a black "snake" of carbon that grows out the top of the beaker.

(1) **Note**: This demo must be used with an in-bench hood. There is a smaller scale, modified version for rooms without hoods

8) Group 17/VII: The Halogens
   a) **Halogens** – Display flasks containing the halogens chlorine, bromine, and iodine.
      i) On request, you can order a special flask of bromine that can be frozen in liquid nitrogen
   b) Display some common products containing halogens, such as a box of iodized salt (NaCl, NaI), a bottle of laundry bleach (NaOCl), a can of freon (CF_2Cl_2), etc.
   c) Interhalogen Compounds.
      i) Show ball-and-stick models of ClF, BrF_3, IF_5
   d) Oxyacids and Oxyanions.
      i) Show ball-and-stick models of the oxoacids of chlorine
      ii) **Sound** – Tap a very small pile of red phosphorus and $KClO_3$ with a hammer to show a reaction that produces light, sound, and heat, and recreates on a larger scale the reaction that occurs when you strike a match.
         (1) This demo tends to send off sparks; you may want to wear a lab coat

9) Group 18/VIII: The Noble Gases
   a) **Gas Discharge Tubes of the Noble Gases** – show that different gases give different colors when subjected to an electric discharge (H_2 tube also available)
   b) Show ball-and-stick models of XeF_2, XeF_4, $XeOF_4$, and XeO_3

Transition Elements

1) Transition Metals
   a) Pass samples of Cr, Mn, Fe, Co, Ni, Cu, Zn, and Cd around the class.
   b) **Multiple Oxidation States of Vanadium** – Shake a solution of ammonium meta-vanadate with a Zn-Hg amalgam to reduce the vanadium from +5 to +4 to +3 to +2 with different colors at each stage
   c) **Multiple Oxidation States of Manganese** – Starting with $KMnO_4(aq)$ in four beakers, carry out reactions to display manganese in the +7, +4, +3, and +2 oxidation states
   d) **Paramagnetic and Diamagnetic Salts** – Provide experimental evidence of electron spin by bringing a powerful magnet close to suspended test tubes of $MnSO_4$, $FeSO_4$, $NiSO_4$, and $ZnSO_4$ to show the different responses due to different numbers of unpaired electrons
   e) **Ferromagnetism** – Demonstrate ferromagnetism by inserting a cow magnet into a test tube inside a glass bottle containing iron filings – the resulting arrangement of the iron filings shows the three-dimensional shape of the magnetic field
   f) **Ferrofluid** – Display the properties of ferrofluid, a magnetic suspension that produces spikes and other shapes when exposed to a magnetic field.

2) Chemistry of Selected Transition Metals
   a) **Ammonium Dichromate Volcano** – Reduce chromium from +6 to +3 by igniting an $(NH_4)_2Cr_2O_7$ volcano in an aquarium to produce N_2 (g), H_2O (g), and Cr_2O_3 (s)
   b) **Chromate-Dichromate Equilibrium** – Show the pH dependence of the $CrO_4^{2-}/Cr_2O_7^{2-}$ system
   c) **Penny vs. Nitric Acid** – Read Ira Remson's own account of his first encounter with chemistry as you repeat his experiment by pouring concentrated nitric acid over a copper penny.

3) Coordination Complexes
   a) Show models with different coordination numbers and geometries
      i) Square planar - cis and trans MA_2B_2
      ii) Octahedral - cis and trans MA_4B_2
      iii) Octahedral - cis and trans MA_2B_2, where A is bidentate
      iv) Octahedral - optical isomers - cis MA_2B_2, where A is bidentate
      v) Octahedral - optical isomers – MA_3, where A is bidentate
   b) **Cobalt Complexes and Temperature v2.0** – Show the color change observed when octahedral $[Co(H_2O)_6]^{2+}$ is converted to tetrahedral $[CoCl_4]^{2-}$
   c) **Complex Ion Formation Ni and Cu**– Show color change associated with formation of complex ions. In a tall graduated cylinder of Cu^{2+} or Ni^{2+} aqueous solutions, carefully add 6 M NH_3 to create a layering effect of $[M(H_2O)_6^{2+}$ / $M(OH)_2$ / $[M(NH_3)_4^{2+}]$
      i) $CuSO_4$ gives a blue / white / dark blue layering
      ii) $Ni(NO_3)_2$ gives a green / white / dark blue layering

4) Ligands with More Than One Donor Atom
   a) **Precipitates and Complexes of Nickel** – Add different amounts of ethylenediamine to beakers of Ni^{2+} to contrast the colors of $Ni(H_2O)_6{}^{2+}$ and the Ni^{2+} chelate complexes with one, two, and three ethylenediamine molecules.

5) Isomerism
   a) Geometrical isomerism
      i) See models above
   b) Optical isomerism
      i) Show pairs of MA_2B_2 or MA_3 enantiomers, where A is bidentate; use a large mirror with the models to help explain the concept of non-superimposable mirror images
      ii) **Polarizing Filters and Limonene** – Place small beakers of (*R*)-(+)-limonene and (*S*)-(-)-limonene between two polaroid sheets on the overhead projector to show the equal but opposite rotation of plane-polarized light by these enantiomers; you can also show that a racemic mixture does not rotate polarized light

6) Color and Magnetism
   a) **Nickel and Cobalt Ammine Complexes** – show the dependence of color on both the metal ion and its oxidation state. Add concentrated ammonia to Ni^{2+} and Co^{2+} solutions to show different colors with the same ligand. Next, shake some of the resulting $[Co(NH_3)_6]^{2+}$ complex with O_2 to shift the oxidation state from Co^{2+} to Co^{3+}, to show different colors with different oxidation states.
   b) **Paramagnetic and Diamagnetic Salts** – Provide experimental evidence of electron spin by bringing a powerful magnet close to suspended test tubes of $MnSO_4$, $FeSO_4$, $NiSO_4$, and $ZnSO_4$ to show the different responses due to different numbers of unpaired electrons

Organic Chemicals and Polymers

1) Organic Compounds
   a) Show ball-and-stick models of butane and cyclobutane to highlight carbon's ability to form chains or rings of atoms

2) Bonding Characteristics of the Carbon Atom
   a) Display a lucite tetrahedron containing a model of methane to illustrate the tetrahedral geometry of carbon in CH_4
   b) Use models to show *sp*^{3}, *sp*^{2}, and *sp* geometry around a central carbon atom
   c) Bring pairs of orbitals close together (such as s and s, s and p, two p orbitals end-to-end, and finally two parallel p orbitals) to show how σ and π bonds are formed
   d) Contrast orbital overlap models and ball-and-stick models of C_2H_6, C_2H_4, and C_2H_2 to illustrate the hybrid orbitals and atomic orbitals used to form the various sigma and pi bonds

3) Saturated Hydrocarbons
   a) Alkanes
      i) Normal Alkanes – start with a model of CH_4 and add -CH_2- units to show the build-up of a homologous series
      ii) Show models of normal and branched isomers (i.e. *n*-pentane, isopentane, and neopentane)
      iii) Classification of Carbon Atoms – use models of ethane, propane, and 2-methylpropane to point out primary, secondary, and tertiary carbon atoms
   b) Structural Formulas
      i) Show ball-and-stick models of simple molecules such as H_2, H_2O, H_2O_2, O_2, O_3, CO_2, CO, CH_4, C_2H_2, C_6H_6, and C_2H_5OH; you may also wish to show models of some elements such as Cl_2, P_4, and S_8. (Please indicate desired models upon ordering.)
      ii) Show an extended lattice model of NaCl to remind students of the structural differences between ionic and covalent compounds
   c) Alkane Isomerism
      i) Contrast models of isomers:
         (1) Butane and isobutene
         (2) Butanol and isobutanol
         (3) Ethanol and dimethyl ether
         (4) 1-chloropropane and 2-chloropropane
         (5) 1,1-dichloroethane and 1,2-dichloroethane
      ii) Fischer Projections
         (1) Use a model of CHXYZ (or CHBrClF) as a visual aid in writing its Fischer projection formula; the enantiomer can also be provided to show that its Fischer projection formula is different.
            (a) Optional: a "flattened" model of CHXYZ can also be provided to represent the Fischer projection.

iii) Newman Projections
  (1) Use the Newman projection device to show eclipsed and staggered views of a simple alkane

d) Cycloalkanes
  i) Show Darling models of cyclopropane, cyclobutane, cyclopentane, and cyclohexane
  ii) Display one (or two) extra large Darling model of cyclohexane and define the terms axial and equatorial
  iii) Show two models of cyclohexane to contrast the chair and boat conformations
  iv) Pass around Darling models of cyclohexane so students can appreciate its unique structure and contrast the chair and boat conformations for themselves
  v) Add colored substituents to two models of a cycloalkane to show *cis* and *trans* isomers

e) Physical Properties of Alkanes and Cycloalkanes
  i) **Miscible vs. Immiscible Liquids** – Mix ethanol and colored water in one beaker and hexane and colored water in another to demonstrate miscibility and immiscibility due to differences in the intermolecular forces of alcohols as the size of the alkyl group increases. (The demo can be repeated with octanol and water, to show that, as carbon chain length increases, polarity of alcohols decreases.)
  ii) Use two (small) space-filling models of pentane and two (small) space-filling models of neopentane to demonstrate how branching decreases London forces with a resulting decrease in boiling point

f) Chemical Properties of Alkanes and Cycloalkanes
  i) **Combustion of Methane Bubbles** – Ignite large soap bubbles filled with CH_4 – this gets a "wow!" even from jaded college students! If a more intense demo is needed, try Methane Bubbles XTREME.
  ii) **Methane Bubbles XTREME** – Ignite a tower of methane-filled soap bubbles to produce a pillar of flame 3-5m high. As seen on Mythbusters!

4) Unsaturated Hydrocarbons
  a) Contrast ball-and-stick models of ethylene and acetylene with the orbital overlap models
  b) Constitutional Isomerism in Alkenes
    i) Contrast models of 1,2-dichloroethane, *cis*-1,2-dichloroethylene, *trans*-1,2-dichloroethylene, and 1,1-dichloroethylene
    ii) Contrast models of butane, 1-butene, *cis*-2-butene, and *trans*-2-butene
  c) Polymerization of Alkenes: Addition Polymers
    i) **Disappearing Styrofoam Cup** – Make a styrofoam (expanded polystyrene) cup disappear by placing it in a dish of acetone
    ii) **Disposable Diaper Demo** – See how much water you can add to a super-absorbent disposable diaper, then cut open another diaper to show the super-absorbent powder, Water Lock J-550, which is polysodium acrylate cross-linked with starch; the original polymer results from multiple addition reactions of the alkene functional groups of acrylic acid molecules. (One diaper holds up to 1 L of water!)
  d) Aromatic Hydrocarbons

i) Contrast two styrofoam models of benzene, one showing *sp*^{2} orbitals and *p* orbitals and one model showing delocalization over the whole ring, and a ball-and-stick model showing three π bonds
ii) Pass Darling models of benzene around the class so the students can observe its symmetrical planar structure and *sp*^{2} bond angles
iii) Add colored substituents to three Darling models of benzene to show ortho, meta, and para substitution

5) Alcohols, Phenols, and Ethers
   a) Bonding Characteristics of Oxygen Atoms in Organic Compounds
      i) Show ball-and-stick models of water, methanol, phenol, and dimethyl ether
   b) Nomenclature for Alcohols
      i) Show models of 1-propanol, 2-propanol, and 2-methyl-2-propanol when discussing 1°, 2°, and 3° alcohols
   c) Physical Properties of Alcohols: Hydrogen Bonding
      i) **Negative Volume of Mixing** – Mix ethanol and colored water in a graduated cylinder to demonstrate their miscibility and negative volume of mixing due to hydrogen bonding.
      ii) **Vapor Pressure: Evaporation** – Show the effect of intermolecular forces on vapor pressure by letting three students make streaks of water, methanol, and acetone on the blackboard and then observing the relative rates of evaporation.
      iii) **Miscible vs. Immiscible Liquids** – Demonstrate miscibility and immiscibility by combining ethanol and colored water in one beaker and hexane and colored water in another beaker – solubility decreases as the size of the alkyl group increases. The demo can also be repeated with octanol and water to show that as carbon chain length increases, polarity of alcohols decreases.
   d) Oxidation of Alcohols
      i) **Money to Burn** – Soak a dollar bill in a water-alcohol mixture and then light it with a match; the high specific heat of water keeps the combustion temperature low enough to prevent burning the bill
      ii) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the flammability of organic vapors
      iii) **Oxidation of Alcohols** – Demonstrate the oxidation of ethanol with $K_2Cr_2O_7$ on the overhead projector; the alcohol solution changes from orange to green to blue as the Cr(VI) is reduced; this reaction is the basis for the Breathalyzer test
   e) Ethers
      i) **Ether Fire** – Demonstrate the combustion of ether by allowing the vapor to flow from a can down an inclined trough to a candle, resulting in a vapor flashback fire.
   f) **Slime!** – make a cross-linked gel by mixing solutions of polyvinyl alcohol and borax; use this demo to relate concepts such as polymers and hydrogen-bonding to a commercial product students are familiar with.

6) Aldehydes and ketones
   a) Show models of formaldehyde, acetaldehyde, and acetone
   b) Physical Properties of Aldehydes and Ketones
      i) Pass around scent samples of various aldehydes and ketones for students to smell, and show the structures so the class can identify the functional groups. Choose from: cinnamaldehyde (cinnamon), benzaldehyde (almond), vanillin (vanilla), (S)-(+)-carvone (caraway), (R)-(-)-carvone (spearmint).
   c) Oxidation and Reduction of Aldehydes and Ketones
      i) **Tollen's Test: Silver Mirror** – Create a silver coating inside a small Erlenmeyer flask using Tollen's reagent and an aldehyde solution.
      ii) **Benedict's Test** – Use Benedict's solution and dextrose to demonstrate Benedict's test for aldehydes. Different sugars available upon request.

7) Carboxylic Acids, Esters, and Other Acid Derivatives
   a) Structure of carboxylic acids and their derivatives
      i) Show models of formic acid and acetic acid
      ii) **Production and Combustion of CO** – Add concentrated H_2SO_4 to formic acid to produce H_2O and CO(g), then ignite the CO(g) to show that it burns with a blue flame.
         (1) Optional: you can also show that, at extremely low pH, phenolphthalein turns orange
   b) Esters
      i) Contrast models of acetic acid and methyl acetate
      ii) Pour a small amount of amyl acetate (or another ester) into a dish of warm water to vaporize the samples so students throughout the room can smell them
      iii) **Dancing Green Goblin** – Burning the methyl ester of boric acid produces a pure green flame that seems to "dance"

7) Amines and Amides
   a) Bonding Characteristics of Nitrogen Atoms in Organic Compounds
      i) Show models of methylamine, acetamide, and acetonitrile
   b) Structure and Classification of Amines
      i) Show models of ammonia, methylamine, dimethylamine and trimethylamine when discussing 1°, 2°, and 3° amines
   c) Amides
      i) Show models of formamide and acetamide
      ii) Show models of the two resonance forms of formamide
   d) Formation of Amides
      i) Show models of acetic acid and NH_3, and combine the models to make acetamide
      ii) **Nylon 6-10** – demonstrate the polymerization of hexamethylenediamine with sebacoyl chloride to produce the polyamide, Nylon 6-10

9) Carbohydrates
   a) Chirality: Handedness in Molecules

i) Mirror Images and Chirality
   (1) Demonstrate the concept of "handedness" or chirality with your own hands and then with a pair of large thermal gloves
ii) Molecular Chirality
   (1) Use a large mirror and a pair of enantiomeric models (R- and S- CHBrClF) to explain the concept of non-superimposable mirror images;
   (2) Show two models of CH_4, two models of CH_3Cl, two models of CH_2Cl_2, two models of CH_2ClBr, and an enantiomeric pair of CHBrClF models to contrast superimposable and nonsuperimposable mirror images; you can also point out the planes of symmetry that exist in all the models except CHBrClF.
iii) Optical Activity
   (1) **Polarizing Filters and Limonene** – Place small beakers of (*R*)-(+)-limonene and (*S*)-(-)-limonene between two polaroid sheets on the overhead projector to show the equal but opposite rotation of plane-polarized light by these enantiomers; you can also show that a racemic mixture does not rotate polarized light

b) Designating Handedness Using Fischer Projection Formulas
   i) Use a model of CHXYZ (or CHBrClF) as a visual aid in writing its Fischer projection formula; the enantiomer can also be provided to show that its Fischer projection formula is different.
      (1) Optional: a "flattened" model of CHXYZ can also be provided to represent the Fischer projection.

c) Newman Projections
   i) Use the Newman projection device to show eclipsed and staggered views of a simple alkane

d) Steroisomerism: Enantiomers and Diastereomers
   i) **Diastereomers (Molecular Structures)** – Show models of a pair of enantiomers with two stereogenic centers, then show models of a second pair of enantiomers which are diastereomers of the first pair

e) Properties of Enantiomers
   i) Pass scent samples of (*S*)-(+)-carvone (odor of caraway) and (*R*)-(-)-carvone (odor of spearmint) around the class so students can experience the dramatic difference in the odors of these enantiomers

f) Classification of Monosaccharides
   i) Show a (permanent) flexible open-chain models of glucose and frustose

g) Cyclic Forms of Monosaccharides
   i) Contrast flexible open-chain models of glucose and fructose with large plastic Darling models of the cyclic forms of □-D-glucopyranose and □-D-fructofuranose

h) Reactions of Monosaccharides
   i) **Combustion of Candy**– Contrast the rate of oxidation of sucrose in the body (by eating some candy) with the oxidation of sucrose by $KClO_3$ (as shown by dropping some candy into molten $KClO_3$, producing steam and a lavender flame. Body temperature is ~37C, and the melting point of $KClO_3$ is 368C.

ii) **Benedict's Test** – Use Benedict's solution and dextrose to demonstrate Benedict's test for aldehydes. Different sugars available upon request.

i) Disaccharides

i) Display a large plastic Darling model of sucrose

ii) **Black Carbon Snake** – add concentrated H_2SO_4 to sugar in a tall beaker to produce a black "snake" of carbon that grows out the top of the beaker.

(1) **Note**: this demo must be used with an in-bench hood. There is a smaller scale, modified version for rooms without hoods

10) Lipids

a) Terpenes

i) pass scent samples of limonene (both enantiomers are found in citrus oils), (*S*)-(+)-carvone (oil of caraway), and (*R*)-(-)-carvone (oil of spearmint) around the class so students can appreciate some of the distinctive odors of terpenes

b) Biological Membranes

i) **Surface Tension** – Demonstrate the disruption of the surface tension of water by detergent by sprinkling baby powder in a large dish of water and then touching the water lightly with a wood stick dipped in detergent. Use the overhead projector or document camera to display the settling of powder through the water.

11) Proteins

a) The Biological Role of Proteins

i) To show the function of the structural protein collagen, contrast a regular turkey drumstick with a rehydrated "rubberized" turkey drumstick from which $Ca_3(PO_4)_2$ and other minerals have been dissolved.

b) Amino Acids: The Building Blocks of Proteins

i) show models of simple amino acids: glycine and alanine

c) Structure of Proteins

i) Display a telephone cord and point out various features of its configuration to help students understand by analogy the terms *primary*, *secondary*, and *tertiary* structure of proteins

12) Enzymes and Vitamins- Biological Catalysts

a) **Catalysis of a Reaction** – Remind students of the function of a catalysis by demonstrating the catalysis of the H_2O_2 decomposition of NaK-tartrate with Co^{2+}. Adding Co^{2+} turns the solution pink, but the solution quickly turns dark green as it begins to react vigorously. At the end of the reaction, the pink color is restored showing regeneration of the catalyst;

i) Optional: Use the solution containing the regenerated catalyst to catalyze the same reaction in a second beaker; this reaction is less vigorous

b) **Catalytic Decomposition of H_2O_2** – Compare and contrast several different catalysts, including chicken liver, used for the decomposition of hydrogen peroxide

i) **NOTE**: this demo requires 48 hours notice

13) Nucleic Acids
   a) The Double Helix of DNA
      i) Show a model of DNA

Seasonal Demos

1) **Beat Michigan** – Pour clear solutions together to create Michigan's colors Blue and Yellow. Pour a second set of clear solutions in and destroy Michigan to make OSU's Scarlet and Grey
2) **Goblet of Fire** – Halloween – Burning the methyl ester of boric acid produces a pure green flame that seems to "dance"
3) **Exploding Pumpkin** – Halloween – React CaC_2 and water to produce acetylene, then light the gas produced to force the pieces out of a pre-carved Jack-o-lantern
4) **Genie in a Bottle** – Halloween – Use MnO_2 to catalyze the decomposition of 30% H_2O_2, producing a large cloud of hot water vapor. The heat generated is intense enough to shrink the 2 L bottle used for the demo
5) **Red, White, and Blue Density Tower** – Fourth of July – Display a large graduated cylinder with layers of red, clear, and blue liquids, (water, mineral oil, and ethanol) and have the students propose explanations for the layering and separation.
6) **Vanishing Valentine** – Valentine's Day – Shake a colorless solution to produce a pink solution for Valentine's Day

Outreach Demos

1) **Show**
   a) **Boiling Water at Room Temperature** – Show water boiling at room temperature in a beaker in an evacuated bell jar, then put your hand in the water after boiling to convince students of its low temperature.
   b) **Charles' Law** – Pour liquid nitrogen over a balloon to show that a decrease in T is accompanied by a decrease in V.
   c) **Combustion of Ethanol Vapors** – Allow a small amount of ethanol to vaporize in a large carboy, pour out the excess liquid, and hold a lighted splint to the mouth of the container – the impressive reaction also demonstrates the explosive flammability of organic vapors.
      i) This one is especially good if you bring the demo in a gallon size water/milk jug and then pull out the larger version.
   d) **Conductivity Testers**
      i) **Sugar and Salt** – Use two conductivity testers with light bulbs to contrast the conductivity of d-H_2O, sugar solution, and NaCl (aq), and tap water
      ii) **Strong and Weak Acids and Bases** – Contrast the extent of ionization in weak and strong acids and bases using the lightbulb conductivity apparatus.
   e) **Elephant Toothpaste** – Demonstrate the decomposition of 30% H_2O_2 in the presence of dishwashing liquid and KI, producing an upsurge of steaming foam.
   f) **He and SF_6 balloons** – Have adults inhale He or SF_6 from balloons and then talk. HI-larious
   g) **Marshmallow Man** – Demonstrate the effect a decrease in *P* has on *V* by placing a marshmallow snowman in a bell jar and then evacuate the jar.
   h) **Pressure Bar** – Allow a few students to experience what 1 atm (14.7 psi) "feels" like by resting an iron bar one inch square and 54 inches long on their toes
   i) **Rainbow Cups** – Add a colorless liquid to 6 "empty" beakers, producing the colors of the rainbow – use this demonstration to show how evaluation of observations and experimental results leads to hypotheses and further testing (the scientific method)
   j) **Yamada and Dry Ice** – Add a chunk of dry ice to a 2 L cylinder containing a basic solution and Yamada universal indicator; the dry ice gradually acidifies the solution causing the color to change in the order purple, blue, green, yellow, orange

2) **Hands-on**
   a) **Alka Seltzer Poppers** – React an Alka Seltzer tablet with water in a film canister (placed inside a tube with a ping pong ball on top) to show that the production of gas is so great it pops the top off the film canister and launches the ball
   b) **Anodizing Titanium Rings** – Use 9V batteries and an electrolyte solution to anodize titanium rings, depositing a layer of titanium oxide and changing the color of the ring
   c) **Baby Volcanoes** – Help kids make baby volcanoes using baking soda, vinegar, and food coloring. Bite the tip off of sugar cones and place over tiny Erlenmeyer flasks to make them look more volcano-like
   d) **Cat's Meow** – Have students investigate the effects of soap on fats by having them apply food coloring to small dishes of milk, then add dishsoap with a toothpick
   e) **Gooyuk/Ooblak** – Pass around bowls of a cornstarch-water mixture; the properties of this non-Newtonian fluid challenge our traditional definitions of liquid and solid (and is awesome).
   f) **Density of Coke vs. Diet Coke** – Drop unopened cans of regular Coke and Diet Coke into an aquarium filled with water. Coke sinks and Diet Coke floats – challenge the class to postulate an explanation.
   g) **Hot and Cold Packs** – Demonstrate dramatic differences in heats of solution by dissolving NH_4NO_3(s) in water in a Ziploc bag to make an instant "cold pack" and $CaCl_2$(s) in water to make an instant "hot pack", then pass the bags around the class
   h) **LN2 Ice Cream** – Make ice cream using a milk/cream base and liquid nitrogen
   i) **Slime** – Mix Elmer's glue and a borax solution (with food coloring) to make an easy cross-linked polymer.
      i) Can be done in large batches or individually
   j) **Pen Chromatography** – Have students make designs or dots on filter paper then use isopropanol, acetone, or ethanol to separate the colors in the pens and markers

Note: This listing was updated by Angela M. Miller to serve as a master list of available demonstrations for general chemistry. It was prepared from the individual class lists to accompany the textbooks. Original lists prepared by Mary H. Bailey.