

# SCOPE, SEQUENCE, and COORDINATION

A National Curriculum Project for High School Science Education

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# SCOPE, SEQUENCE, and COORDINATION

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**National Science Education Standard—Physical Science  
Structure and Properties of Matter**

Matter, as found in nature, consists primarily of mixtures, compounds, and elements in various proportions. The observable properties of mixtures depend upon the nature of the components. A mixture can be separated into pure substances using the characteristic properties of the substances contained in the mixture.

## Teacher Materials

Learning Sequence Item:

# 1024

## Properties of Aqueous Solutions

*May 1996*

*Adapted by: George Miller*

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**Mixtures, Elements, and Compounds.** Students should measure properties of aqueous solutions, such as boiling point and freezing point, and deduce that they vary in a predictable way with composition and concentration. The phenomenon of diffusion should be investigated using easily distinguished solutes to establish that exchange occurs as a result of concentration gradients (*Chemistry, A Framework for High School Science Education, p. 51*).

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1. Is It Hot Enough to Melt?
2. Melting Ice is a Good Idea!
3. Boil This!
4. Diffusion in Every Phase

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# 1024

**Mixtures, Elements, and Compounds.** Students should measure properties of aqueous solutions, such as boiling point and freezing point, and deduce that they vary in a predictable way with composition and concentration. The phenomenon of diffusion should be investigated using easily distinguished solutes to establish that exchange occurs as a result of concentration gradients (*Chemistry, A Framework for High School Science Education, p. 51*).

## Learning Sequence

Science as Inquiry	Science and Technology	Science in Personal and Social Perspectives	History and Nature of Science
<p>Is It Hot Enough to Melt? <b>Activity 1</b></p> <p>Melting Ice Is a Good Idea! <b>Activity 2</b></p> <p>Boil This! <b>Activity 3</b></p> <p>Diffusion in Every Phase <b>Activity 4</b></p>	<p>Melting Ice Is a Good Idea! <b>Activity 2</b></p> <p>Boil This! <b>Activity 3</b></p> <p>Winter Roads <b>Reading 1</b></p>	<p>Melting Ice Is a Good Idea! <b>Activity 2</b></p> <p>Boil This! <b>Activity 3</b></p>	

# Suggested Sequence of Events

## Event #1

### Lab Activity

1. Is It Hot Enough to Melt? (45 minutes)

## Event #2

### Lab Activity

2. Melting Ice Is a Good Idea! (45 minutes)

## Event #3

### Lab Activity

3. Boil This! (45 minutes)

## Event #4

### Lab Activity

4. Diffusion in Every Phase (45 minutes)

## Event #5

### Readings from Science as Inquiry, Science and Technology, Science in Personal and Social Perspectives, and History and Nature of Science

#### The following reading is included with the Student Materials:

#### Reading 1 Winter Roads

#### Suggested additional readings:

- Better Snow and Ice Control: Using State-of-the-Art Technologies*, U.S. Dept. of Transportation, Federal Highway Admin. Announcement, Washington, D. C.: U. S. Govt. Pubs. TD2.2:SN6, 1995–96, pp. 4–7.
- Bohren, Craig, “Melting with Salt and Heating with Ice.” *Weatherwise*, Vol. 46, No. 6, December 1993/January 1994, pp. 46–48.
- Coghian, Andy, “Organ Transplants Prepare for the Big Chill.” *New Scientist*, July 1993, p. 17.
- Cossins, A. R., and K. Bowler, *Temperature Biology of Animals*, London: Chapman and Hall, 1987 (Internet).
- Chemical Deicers and the Environment*, Chapter 23, D’Itri, Frank M., Ed., Ann Arbor: Lewis Publishers, 1992, pp. 553–557.
- Davenport, J., *Animal Life at Low Temperature*, London: Chapman and Hall, 1992.
- “Frozen Insects Will Thaw into Summer Pests,” *Lincoln Journal Star*, March 3, 1996.
- Mark, F. W., and W. Jetten, “Propylene Glycol, A New Base Fluid for Automotive Coolants,” *Engine Coolant Testing: Second Symposium* ASTM STP 887, 1986, p. 61.
- Potapov, Eugene, “How Salamanders Survive the Deep Freeze.” *New Scientist*, Sept. 11, 1993, p. 15.
- Selection and Use of Engine Coolants and Cooling System Chemicals* ASTM STP 120B, 1974, pp. 1–3.
- Shojai, Amy D., “Health Handbook,” *CATS Magazine*, Inc., 1995 (Internet).

*Assessment items are at the back of this volume.*

## **Assessment Recommendations**

This teacher materials packet contains a few items suggested for classroom assessment. Often, three types of items are included. Some have been tested and reviewed, but not all.

1. Multiple choice questions accompanied by short essays, called justification, that allow teachers to find out if students really understand their selections on the multiple choice.
2. Open-ended questions asking for essay responses.
3. Suggestions for performance tasks, usually including laboratory work, questions to be answered, data to be graphed and processed, and inferences to be made. Some tasks include proposals for student design of such tasks. These may sometimes closely resemble a good laboratory task, since the best types of laboratories are assessing student skills and performance at all times. Special assessment tasks will not be needed if measures such as questions, tabulations, graphs, calculations, etc., are incorporated into regular lab activities.

Teachers are encouraged to make changes in these items to suit their own classroom situations and to develop further items of their own, hopefully finding inspiration in the models we have provided. We hope you may consider adding your best items to our pool. We also will be very pleased to hear of proposed revisions to our items when you think they are needed.

## Science as Inquiry

**Is It Hot Enough to Melt?****What is the melting point of different substances?****Overview:**

This activity introduces students to the technique of measuring the melting (freezing) temperature of micro-scale amounts of substances and has them examine the effect of forming solutions by mixing substances. It is suggested as a starting point for this unit, even though it does not involve aqueous solutions, because it can be done using water above room temperature as the heating fluid. This avoids ice or ice-salt bath combinations which will need to be used in subsequent activities to obtain the sub zero temperatures needed to study the effects of dissolved substances on the freezing temperature of water. A few of the suggested combinations involve melting temperatures above that of boiling water. Pure ethylene glycol (B.Pt. 198°C) or mineral oil (use up to 230°C ) can be used as the bath fluid if such combinations are preferred. Note that only a few groups might need to go above 100°C.

These methods are suggested to improve safety, the expense of use of large amounts of chemicals, and to reduce expensive and hazardous waste disposal. Students will need time to practice and become accustomed to working at the micro (or semi-micro) level. It is also instructive for students to experience that not all solutions involve water as a solvent.

**Materials:****Per lab group:**

chemicals, any 2 (only a few g per class)

phenyl salicylate (salol) (43)\*

lauric (dodecanoic) acid (44)

hydrocinnamic acid (48)

benzophenone (49)

stearic acid (70)

durene (79)

naphthalene (80)

glutamic acid (98)

resorcinol (110)

benzoic acid (122)

malonic acid (136)

maleic acid (140)

hydroquinone(173)

melting point setup

capillary melting-point tubes, 2

microscope slides (or small watch glasses), 2

spatula (or firm drink stirrer)

\*Number in parenthesis is melting temperature of chemical in °C.

**Procedure:**

Students work in pairs for this activity. One-half of the class measures the melting temperature of pure substance A, the other half measures the melting temperature of pure substance B. The pooled class

data should be examined for statistical variation around each of the two temperatures and the ideas of distributions, means and precision can be reviewed.

Each pair of students then undertakes to measure a specific mixture of A and B so that the class covers the whole range from (5% A, 95% B) to (5% B, 95% A). The class should plot the pooled results, and compare observations. The teacher can either pre-prepare the mixtures by weighing small quantities into storage vials, or may have students weigh out their own samples, depending on the number and type of balance available.

### Background:

For specific information on melting procedures, see the Technique Sheet “How to Determine the Melting Point.” Similar types of compounds (e.g., two acids, or two aromatic (benzene ring containing) substances) are likely to be better combinations. Substances listed are selected based on availability in a common high school supply catalogue. Depending on what is available to the teacher, select low melting substances if desired to use only water baths. Not all combinations have been tried. Teachers and students are invited to submit results to the program. Many other possibilities exist, but are less readily available.

While it is not the objective of this unit to focus on energy transfer, it will be an excellent time to review this for the melting/freezing transition (see “General Background Material” below).

Pure substances undergo melting at a very specific temperatures and pressures, since it takes a definite amount of energy to overcome intermolecular forces in the solid phase and permit the type of molecular motion needed to form the liquid. At the melting point (probably more correctly called the melting temperature) the substance is transitioning from the solid to the liquid phase. A specific amount of energy called the latent heat of fusion is absorbed in this process for a standard mass (or number of moles) of the substance. Until all this energy is absorbed and distributed among all of the molecules of substance present, the temperature of the substance cannot increase. When the opposite transition process, freezing, takes place the same effect is observed, with the same amount of latent heat being released to the surroundings or to the cooling agent. When substantial amounts of substance are present, there may

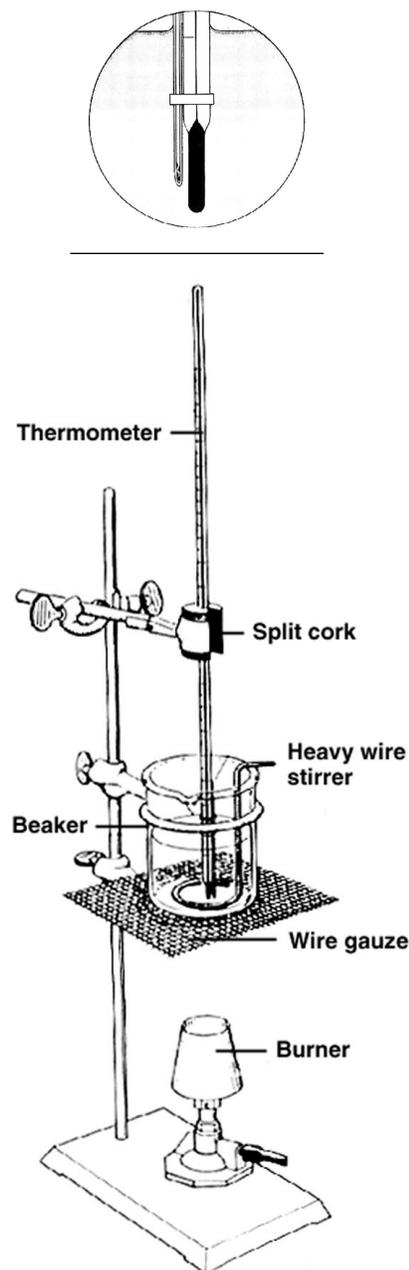


Fig. 1. Melting Point Setup

be some time effects in energy absorption/ release leading to the phenomena of super heating or supercooling. These can be examined using “cooling curve” experiments. Students should probably carry out a cooling curve experiment, if they have never done one, prior to doing melting points determinations in order that they appreciate the transitions involved (see, for example, “Lab 18: Heating and Cooling Curves,” *Chemistry*, Prentice Hall, M. Wagner).

Students should be sure to powder any crystals carefully and tap them gently into the capillary tubes. They will need to note the temperature at which they observe crystals start to melt, since there will be a wide range of temperature over which the non-eutectic mixtures will be melting. They then will observe when all the crystals seem to have melted. If time permits, they can measure the freezing points by allowing cooling to occur.

**Expected Results.** Typical results obtained for a specific pair of substances (camphor and hydroquinone) are shown in Fig. 2 (see also Table 2 in “Variations”). Not all pairs will give similar results, but many will. Teachers who encourage students to experiment with the above combinations may expect that occasionally more than one eutectic can form, or even more rarely, none at all.

**General Background Material.** Solutions may be defined as any homogeneous mixture of two or more pure substances. Solutions can be formed from mixtures of substances of one or more phases: gas-gas, gas-liquid, gas-solid, liquid-liquid, solid-liquid, and solid-solid are usually listed. Once a solution is formed, by definition it must form only a homogenous single phase. The most familiar solutions are those formed by dissolving solids in liquids such as water. In most cases one substance is in excess quantity and is called the solvent, then any other substances are called solutes. Understanding solutions is generally difficult for students, even to quite advanced levels. The aim in this unit is to advance investigation and discussion so that understanding progresses. The unit also emphasizes how important solutions are in the everyday world. Understanding of solutions and the solution process is not expected to be anywhere near complete. Frequent classroom open discussions and exchange of ideas are to be encouraged throughout with every consideration given to exploring alternative models, rather than simple enforcement of the “right” one. Students often do not recognize air as a common solution. Since they have studied gases rather extensively, recognizing air and other gas mixtures as solutions, may help them understand liquid solutions more easily by analogy. At the end, part of the final activity suggests that they be encouraged to recognize that solid solutions may exist.

In a recent article in *Science Education*, (Vol. 79 (1) pp. 1–17, 1995, John Wiley and Sons, Inc.), entitled

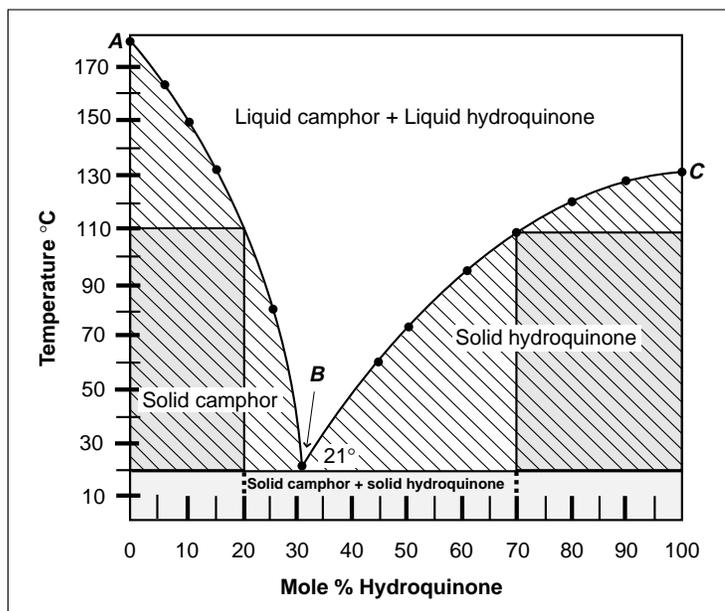


Fig. 2

“Relational Conceptual Change in Solution Chemistry,” Jazlin V. Ebenezer, and P. James Gaskell discuss students ideas about solutions and the difficulty of changing their conceptual views. Part of the need identified is to challenge ambiguities introduced by everyday use of terms such as melting, particle, and solution, and have students learn to appreciate the more precise uses of these terms in the science community.

**The Kinetic Molecular Model of Solutions.** The most frequent solution process encountered by students is solid dissolving in liquid, such as sugar in lemonade, or salt in water. Research studies have shown that many students confuse the solution process with melting (“sugar melts when added to hot tea”) for example. This may be because they are confusing macroscopic observations (the sugar granules seem to melt) with the microscopic phenomenon that is taking place when individual sugar molecules break from the sugar surface by forming stronger linkages with water molecules (hydrogen bonds between the OH functional groups on the sugar molecules and the OH of the water) so that each sugar molecule becomes separately surrounded by water molecules.

The same is happening when an ionic crystal such as sodium chloride (common salt) dissolves in water. The resulting solution, consisting of positive sodium ions surrounded by water molecules oriented with negative oxygen closest to the sodium ion, and negative chloride ions surrounded by water molecules with the positive hydrogens oriented inward to the chloride ion, is more stable than the original solid - water separated arrangement and so the salt dissolves. Each ion with its water sheath can almost be thought of in the same terms as a “neutral” molecule since attraction or repulsion between these “water-sheathed ions” is greatly reduced over that between the bare ions in the crystal.

By this time in the overall sequence of 9th and 10th grade units the student has had considerable experience experimenting with and thinking about gases as invisible moving particles. It is a natural extension to begin to apply this model to solutions with the difficult step that just because there is now something macroscopic by way of a liquid or solid that is easy to “see” this does not change the kinetic-molecular view that the important “particles” to the model used to explain behavior are still too small to be seen.

Even though this framework statement does not address energy considerations, it is impossible to separate the issue of energy content from the idea of freezing or boiling. Thus earlier units that study phase changes, in particular liquid to gas and liquid to solid, should be revisited for relevancy to this unit, and activities carried out if they have been skipped.

**Melting is Freezing and Boiling is Condensing.** It is important at the outset to ensure that it is fully understood that the melting of a solid to a liquid and the reverse process of freezing of the liquid to the solid are the same process when viewed from the kinetic molecular model of substances. We use different terms to describe the phenomena since one occurs on adding energy (heating) to a solid, the other on removing energy (cooling) from a liquid. Observed differences occur because of non-equilibrium phenomena such as supercooling or superheating involving rates of energy transfer and molecular re-orientation. Similar parallel situations are boiling a liquid and condensing a vapor (gas). Thus from a kinetic molecular model viewpoint, these parallel transitions can be treated as the same phenomena. In practice, therefore, it is possible to observe a melting point on heating a solid substance and a freezing point on cooling the same liquid substance that are at the exact same temperature. If they are not, then the difference has to do with the experimental set-up, usually differences in rates of heating or cooling, and not with the underlying molecular model.

Familiar solution properties such as freezing temperature and boiling temperature (usually known as freezing point and boiling point respectively), vary with concentrations, as well as with external pressure.

The latter will not be considered here - all experiments will be conducted at roughly constant atmospheric pressure.

**Pure Substances.** Starting with a pure liquid substance, freezing occurs when sufficient energy is transferred out of the substance to permit most molecules in the substance to lose most freedom of movement. Cooling has reduced translational kinetic energy until it has become approximately equal to the energy provided by attractive forces between the molecules. Until most molecules in the sample have lost this energy, the temperature of the substance remains constant, and so a specific freezing temperature is observed. Conversely, when boiling, kinetic energy in translational motion is large enough to overcome almost all remaining energy of intermolecular forces, and molecules are freed into the gaseous state. here, too, until all molecules have entered the state, energy input from outside is used in freeing new molecules rather than being applied to those already in the gas. Thus, a specific constant boiling temperature is observed.

**Solutions.** When two or more substances are present, molecules of solute substances “interfere” with molecules of the solvent so that intermolecular attractions are generally weaker. Thus more kinetic energy must be removed in order that the molecules form a solid. The freezing temperature for almost all dilute solutions is thus lower than for the pure solvent. The quantitative amount of freezing point lowering for a particular solvent can be related to the concentration of solute in the original solution. In fact, in agreement with the model just presented, the lowering is related directly to the relative numbers of ions or molecules (solute vs. solvent) present in the original solution.

However, one new phenomenon is observed - the freezing point is often spread out over more of a range than for a pure substance. This is because molecular clusters of the two substances can form within the solution even in previously homogenous solutions. Thus heterogeneous solids are formed which can vary in amounts, and so the apparent freezing (or melting) occurs over a range of temperatures rather than at a single temperature. This phenomenon had, in the past, great utility in confirming the identify of a chemical substance. If a preliminary identity was established, and a pure known sample of the probable substance was available for comparison, then a “mixed melting point” was determined. a small amount of “known” was mixed with the “unknown” and the melting point measured. If a sharp melting temperature was observed, at the same temperature as for the pure substance, then the identity was confirmed.

One more important phenomenon is observed in freezing solutions. For many pairs of substances, there is a specific concentration at which when freezing, a coherent, single solid forms instead of a heterogeneous mixture of solids. Such a substance is known as a Eutectic. The model outlined above would predict that a eutectic would have a well defined melting (or freezing) point, but that is lower than either of the pure substances that went to make up the eutectic solution. This is indeed observed. Important examples of eutectics in the everyday world are solders used for joining metal sheets or parts such as copper pipes in houses. It is an advantage to have a specific low melting metal for such use so that joints can be made without melting the parts themselves. As an example, a mixture of 37% lead (m.pt. 328°C) and 63% tin (m.pt. 232°C) metals forms a eutectic which melts at 183°C. If Activity 1 is done as a class exercise, this explores the realm of eutectics, as does the alternate demonstration suggested.

An entirely parallel phenomenon occurs with boiling solutions, except that in this case, the interference by the solute molecules causes a greater input of energy to be needed to reach the boiling point which results in a higher boiling temperature. The solute molecules are effective in inhibiting the solvent mol-

ecules from leaving the surface and forming gas phase molecules, so that additional translational kinetic energy must be present for the boiling process to take place. Just as with freezing, the quantitative nature of the effect depends upon the nature of the solvent, but also on the relative number of solute molecules present (per given number of solvent molecules). Thus boiling points of most solutions are higher than those of the pure substances that made them up. Some exceptions occur when both solvent and solute are very volatile (low boiling points). Just as with the freezing phenomenon, the solution boils over a range of temperatures, a fact which is exploited in the distillation process, since the vapor composition is not the same as the liquid composition. the boiling equivalent of a eutectic is also known - such solutions are called constant boiling or azeotropic mixtures and have specific concentrations. Common examples are alcohol-water (95% alcohol, 5% water) and hydrochloric acid (20.2% HCl, 79.8% water ) both at 760 mm pressure.

### Variations:

**Activity.** A small scale method may be employed for this activity instead of the micro-method.

**Materials.** 12 test tubes (10-15 mm ID); 12 thermometers (suitable range); chemicals—two selected from the “Materials” list above.

This alternate method uses modest amounts of chemicals, 1–5 grams and may be simpler for certain types of classes with fewer facilities. Weighing the chemicals can be done in advance. The experiment is the same as above, except that test-tubes are pre-prepared with weighed amounts of the two substances (Table 1).

**Table 1.** For a class of 33 students, 11 tubes could be made as follows:

Substance A (g)	Substance B (g)
5	0
4.5	0.5
4	1
3.5	1.5
3	2
2.5	2.5
2	3
1.5	3.5
1	4
0.5	4.5
0	5

These amounts roughly correspond to the points shown on the Fig. 2. Each student then clamps a thermometer in the tube so that it is among the powder but not touching the sides. The setup is then heated in a bath as usual. Each group then reports their results to the class and a class graph is constructed. With care, these tubes can be used year after year. It may be difficult to remove and re-insert the thermometer once the material has been melted. It is easier to retain the thermometer-tube set-up, store carefully, and use as -is each year.

**Demonstration.** Another alternative is simply a demonstration. Pairs of substances are stirred together in a glass petri dish on an overhead projector, or by student groups. Powder each pure substances well before mixing so the mixing occurs easily. When the correct proportions have been reached, the mixture will melt below room temperature. An alternative is to carry out the mixing in a test tube and pass around for observations by students. A thermometer can then be used to measure the temperature at which the

mixture melts. For a good temperature measurement keep the test tube insulated with a little foam sheet wrapped around it.

The following mixtures have been suggested in the literature. Not all are compounds that are as safe or obtainable as those listed in the class experiment. Some of the systems suggested above will also have melting points below room temperature over part of their composition range.

**Table 2.**

<b>Substance A (pure m.p.°C) (pure m.p.°C)</b>	<b>Substance B (pure m.p.°C) (pure m.p.°C)</b>	<b>Proportion by mass for eutectic (A:B)</b>
piperonal (35-37)	benzophenone (49-51)	4:6
phenylsalicylate (salol) (42)	benzophenone (49-51)	4:3
orthonitrophenol (44.5)	alpha-naphthylamine (48)	10:9
orthonitrophenol (44.5)	para-toluidine (37.5)	11:9
alpha-nitronaphthalene (57)	alpha-naphthylamine (48)	1:1
camphor (178)	hydroquinone (131)	3.7:6.3

Adapted from:  
None.

Science as Inquiry/Science and Technology  
Science in Personal and Social Perspectives

## Melting Ice is a Good Idea!

### What influences the freezing point of ice?

#### Overview:

This activity continues the exploration students have begun with techniques of measuring the melting (freezing) temperature of different substances. Before starting this experiment, students need to read “Winter Roads” (Reading 1).

#### Materials:

##### Per lab group:

chemicals (a few g of each per class):

sodium chloride

magnesium chloride

calcium chloride

potassium acetate

calcium magnesium acetate (CMA)

ice

salt (NaCl)

distilled or deionized water

freezing point setup (see illustration)

capillary melting-point tubes, 2–4

microscope slides (or small watch glass)

spatula, small (or firm drink stirrer)

dropping pipette w/small capillary end

hand lens

chemical balance, weighs to 0.10 g

measuring cylinder (or pipette), 10-mL

#### Procedure:

Assign compounds to groups of students. They are to research their compound(s) to try to answer the questions posed. They need to measure freezing points of their pure water, by the micro (or small sample) method practiced in Activity 1, and at least two concentrations of their compound(s) in water.

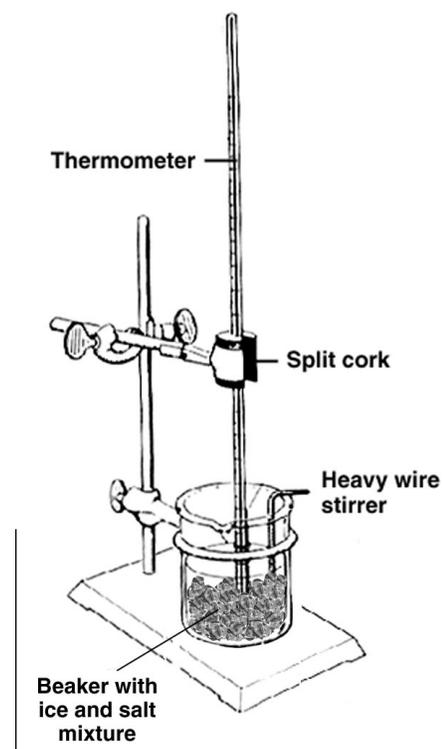
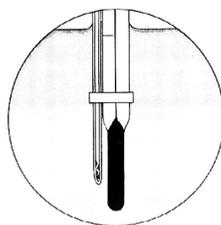


Fig. 1. Freezing Point Setup

**Table 1.** Calculations of the theoretical effect of a 1% (w/w) solution for each substance listed.

Compound	Formula	Molar Mass	Freezing Point of solution 1g in 100g water (calculated)
Sodium chloride	NaCl	58.5	-0.636°C
Magnesium chloride	MgCl <sub>2</sub>	95.22	-0.585°C
Calcium chloride	CaCl <sub>2</sub>	110.99	-0.504°C
Potassium acetate	KOOCCH <sub>3</sub>	98.15	-0.379°C
Calcium magnesium acetate (CMA)	CaMg <sub>2</sub> (CH <sub>3</sub> COO) <sub>6</sub>	442.97	-0.378°C

Solutions suggested are 5% and 10% by mass. If time permits, then in-between measurements can be made (1%, 2%, 7% for example). The cooling is provided by using ice and salt in the cooling bath. Students may have to remove the set-up from the bath briefly to observe, then replace it rapidly so the temperature remains about the same. (See Activity 1 illustration for melting setup.)

### Background:

For specific information on melting procedures, see the Technique Sheet “How to Determine the Melting Point,” which can also be used to find the freezing point. Using tap water will give unusual results since it already contains dissolved substances. If using the small sample method, 100 g of each chemical should be available per class. CMA may be available in some areas as a road de-icing compound. One trade name used was “ICE-B-GON.” If you live in a cold winter area, try contacting your local highway department. If not, it may be prepared by the method described below.

**Calcium-Magnesium Acetate (CMA).** CMA was first produced by reacting lime prepared from the natural mineral dolomite with glacial acetic acid. Dolomite is an approximately 1:1 mixture of calcium and magnesium carbonates (CaCO<sub>3</sub>•MgCO<sub>3</sub>), so the lime prepared by heating this is (CaO•MgO). The resulting CMA was thus equimolar calcium and magnesium acetates. Since that time, commercial research has shown that a more optimal substance for de-icing use is to have a *Ca acetate:Mg acetate* ratio of 7:3. This was developed as the ratio for ICE-B-GON by Chevron Chemical Company (1990). This has been shown to contain mostly crystals of the double salt CaMg<sub>2</sub>(CH<sub>3</sub>COO)<sub>6</sub>.

**Preparation.** CMA may be prepared by mixing approximately 47 g of anhydrous calcium acetate with approximately 100 g of anhydrous magnesium acetate and dissolving in the minimum amount of distilled water (magnesium acetate is very soluble and calcium acetate is not so soluble—so be sure to dissolve all the calcium acetate) and then allowing the solution to evaporate to almost dryness, then drying briefly at 100°C. If the hydrated salts are available, they may be used instead, and the quantities adjusted accordingly. Much of the research in trying to lower the cost of production of CMA revolves around finding a cheaper way to make acetic acid.

Refer to “Background” in Activity 1 to review detailed information on melting, freezing and boiling processes.

### Answers to Student Questions.

These questions may require additional student research to find good answers. An excellent book reference is *Chemical Deicers and the Environment*, by Frank M. D'Itri, Lewis Publishers, Inc., 1992. Some brief answers to the student questions are given below.

1. Yes, the results of the experiments should show that all added substances change the freezing temperature of water, to less than 0°C.

2. It depends on the compound, and on how much of the compound is dissolved in the water. Dissolving twice as much compound generally lowers the FP by twice as much. The actual value should be 1.86°C for every 1 mole of ions or molecules dissolved in 1,000 g of water. For instance, 58.5 g (1 mole) of sodium chloride (NaCl) dissolved in 1 kg of water will have a freezing temperature of -3.72°C. This is calculated thus: 1 mole of NaCl gives 1 mol of Na<sup>+</sup> ions, and 1 mole of Cl<sup>-</sup> ions, for a total of 2 moles of ions in water. Thus the effect = 2 x 1.86°C lowering or -3.72°C since pure water freezes at 0°C. Translating this to a % by mass value means that each 1% by mass of sodium chloride will lower the freezing point by  $10/58.5 \times 2 \times 1.86 = 0.64^\circ\text{C}$ . Theoretical results for the other substances are shown in Table 1. Note: Students at this level are yet not expected to work in molar concentrations.

3. On a mole for mole basis, those compounds that give the largest number of ions in solution will have the largest effect. However, since students at this level haven't learned a lot about moles or ions, they can only compare them directly on a mass basis. Those compounds that have more sub-units in their formula, and those that have the smallest molar mass (composed of atoms with the smaller atomic masses from the periodic table) will have the greater effect since a given mass will contain a greater number of moles of compound dissolved. If students exceed the solubility, and all the substance is not dissolved, then this can also upset predictions. They should be on the lookout for such effects (see Table 1).

4. Adding solute particles (ions, molecules) changes the forces inside the solution, generally weakening them and adding more disorder. Thus more energy must be removed before the high ordered state of a solid can be formed.

5. It is cheap, effective (see above table) , and readily available.

6. Salt-water solution accelerates corrosion of iron and aluminum parts of automobiles, bridges and the iron bars (rebars) used to reinforce concrete roads and structures. In 1987, New York State estimated that for every \$1 spent on road salt, the cost of damage was \$57, in the form of crumbling roads and bridges, corroded motor vehicles, damaged utility lines, and contaminated drinking water. Environmental run-off is hurting plants, especially trees adjacent to highways. Plants cannot survive if there is too much salt in the soil water. This partly has to do with osmotic pressure effects (e.g., pickling process!)

7. Deicing means that you are trying to remove ice that has already formed. Anti-icing means trying to prevent the ice forming in the first place. The latter is more efficient for at least two reasons:

a) Cars and trucks do not have to drive on ice while the spreader works. If the temperature is low, it may take a while for the de-icer to be effective, especially if it is added as a solid, since solid-solid interactions are slow; and b) The anti-icer is spread as a liquid solution which makes it more uniform.

8. Cold air can reach both the top side and the underneath side of roadbed material on a bridge, whereas a regular road is on a bed of "infinite" soil. Thus falling moist air temperature, the most likely cause of road icing, has a quicker impact on bridges than on the regular roads. Thus ice will form on bridges before it begins to form on the roads leading to the bridge. This is so much of a problem, that using more expensive de-icing materials may be justified. See the optional reading: "CMA Use on the Milwaukee Bridge" for an example.

9. These include, costs, availability, ease of applying, ease of storage, possible toxicity to workers, potential for environmental damage.

**Variations:**

Small samples method instead of micro-method. The freezing point of 1–3 mL of solution can easily be measured in a small test tube fitted with a thermometer as suggested in the Variations section of Activity 1. The quantities of chemicals needed are still quite modest, especially if students work in groups.

Adapted from:  
None.

Science as Inquiry/Science and Technology/  
Science in Personal and Social Perspectives

## Boil This!

### How do boiling points for solutions differ?

#### Overview:

In this activity, students determine the boiling points of glycol solutions. Before starting this experiment, students need to read “Salamanders in the Deep Freeze” and “Organ Transplants and the Big Chill,” (Readings 2 and 3).

#### Materials:

##### Per lab group:

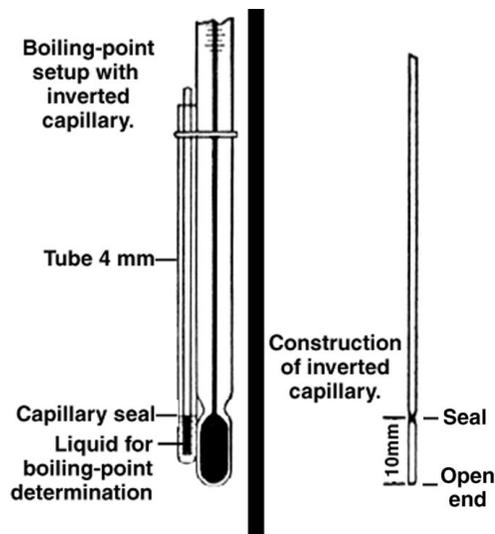
boiling point setup which incorporates:  
test tube, closed end, about 5 mm internal diameter  
melting point capillary  
fine-point dropping pipettes, 2  
distilled or deionized water  
ethylene glycol,  
propylene glycol  
chemical balance

#### Procedure:

Assign student groups specific composition ranges and systems (either ethylene glycol-water or propylene glycol-water). Compositions should range from 100% of one component to 100% of the other component. Each 10% range (10, 20, 30, 40, 50, 60, 70, 80, 90) can be chosen, or each 5% depending on class size and the time assigned to this activity. If a good balance is available then students can make up there few drops of liquid mixtures by mass. If not then they could do it by counting drops. If dropping pipettes have constant size orifices, which good quality ones do, then the drop method works provided that pipettes are held vertically for each delivery. The constant orifice can be maintained by having the whole class use the same pipette to deliver drops into their own tubes from a class storage bottle of the compound or the water.

#### Background:

For specific information on boiling procedures, see the Technique Sheet “How to Determine the



Boiling Point.” If using the micro method, only a few mL of each will be need for a class. This will reduce disposal concerns as well as purchase costs.

Using tap water will give unusual results since it already contains dissolved substances. In a similar manner, use of commercial antifreeze is not recommended as it has additives which will affect the outcome of the experiment.

Refer to “Background” in Activity 1 to review detailed information on melting, freezing and boiling processes.

#### Answers to Student Questions.

1. Yes, the boiling point increases.
2. Yes, the more added, the higher the B.Pt., but it isn't linear, it's a curve.
3. Yes, but the differences are not great between ethylene glycol and propylene glycol.
4. Automobile engines run at high temperatures, but they need a coolant to prevent parts from overheating. In winter, this coolant must be prevented from freezing when the engine is not operating. The higher the boiling point of the mixture, the less fluid will be lost by evaporation or “boiling over” while the engine is operating, especially in very hot climates .
5. The molecules all have hydroxyl (OH) groups. This is important because it means that the compounds will always mix extremely well with water, which also has OH groups.
6. The toxicity of antifreeze is quite high, and it may be harmful to the environment when improperly disposed of. Switzerland has banned the substance. Propylene glycol is less toxic, and appears to be as effective. It's major problem is that it is more expensive.
7. One reason is cost. Water is very inexpensive. Another reason is viscosity. Glycol compounds are quite viscous, and even when hot, when they are much less viscous, may not circulate as freely in the engine as water-glycol mixtures. There is concern that heat removal would not be as effective with pure glycols. Students should have noted the viscosity when trying to measure out samples to measure the boiling points.
8. See the “Background” section in Activity 1 for detailed answer.

#### Variations:

Using larger quantities of materials. If this method is employed, small measuring cylinders may be used for preparing samples (percentages will be on a volume/volume pre-measurement basis. This can be converted by having students measure density of pure substances, and water (or by obtaining this data from Table 1 below) and calculating masses from measured volumes and know densities. Volumes cannot be determined additively, because glycol-water mixtures change volume on mixing.

**Table 1.**

Compound	Density gm/mL at 20°C	Boiling Point 760 mm pressure
Water (H <sub>2</sub> O)	0.9982	100°C
Ethylene Glycol (1,2 Ethanediol), HOCH <sub>2</sub> CH <sub>2</sub> OH	1.1088	198°C
Propylene Glycol (1,2 Propanediol), HOCH <sub>2</sub> CH <sub>2</sub> CH <sub>2</sub> OH	1.0361	189°C

Adapted from:  
None.

## Science as Inquiry

**Diffusion in Every Phase****What are the properties of diffusion?****Overview:**

Student groups are challenged to design a presentation to the rest of the class to illustrate the formation of a solution.

**Materials:**

**Per lab group** (2–10 students):

Erlenmeyer flask

dilute acid solution (or white vinegar) 50 mL

dilute sodium hydroxide solution (0.1 to 1 M)

dropper pipette

dropper bottle with phenolphthalein reagent

potassium permanganate (or crystals of brightly colored, water-soluble chemical)

distilled water

test tube, large, (or graduated cylinder or 1 liter plastic bottle)

sand, brightly colored (one or more colors—aquarium store), 25 mL

sand, beach (about same grain size as colored sand), 500 mL to 1 L

small dropper bottle containing a few drops of inexpensive perfume, ammonia solution, or other compound with strong—but not unpleasant—odor (try to avoid anything to which a student might have a severe allergic reaction)

**Procedure:**

Each group is assigned just one type (gas-gas, liquid-liquid, solid-liquid, or solid-solid) and a competition run for the best presentation. The presentation must illustrate the kinetics of solution formation (i.e., that mixing takes time, and different amounts of time, depending on distances and the phase(s) involved). The presentation must also discuss the model for the structure of matter that supports the ideas illustrated in the presentation.

The teacher may choose to have the class create a rating system, in advance, whereby students establish criteria by which they will judge the effectiveness of other presentations.

**Gas-Gas.** Students are placed at measured distances from one corner of the room as “reporters.” More than one reporter will be at each distance (e.g., each meter at 1–10 meters) in order to sample relative student sensitivity to the odor. Direction might also be important, so all students have a “place” in the measuring grid. The bottle of odor chemical is opened in one corner of the classroom at a specific time noted on a timer or clock. Students report when they first smell the odor. Plots are made of distance/direction vs. time. Discussion might include concerns about the type of smell (perfume might be compared to ammonia, for example), the effects of drafts if someone were to open the classroom door during the exercise, the effects of air conditioning/heating ducts, changes in temperature (compare quantitative results from first class of day to afternoon class, for example).

**Solid -liquid.** A tall, graduated cylinder (or narrow glass bottle) is filled with water. (Permanganate may stain a plastic bottle and obscure some of the effect, though this loss may not be severe.) A crystal of potassium permanganate is dropped to the bottom (or inserted using long forceps to the bottom) as a timer is started. The time that the solution mixes to certain marks on the container are recorded, as is the time when the solution can be considered to be completely homogenous. The ensuing discussion should include comparisons to the gas experiments (noting the vast differences in the distances traveled), and noting effects of stirring because of the insertion process, or the fact that the water had not become completely still before the experiment was begun. This is an important time to point out that a dissolving process is also taking place, as well as a diffusion process. Creative students might be challenged to think how to design the experiment to avoid some of these differences. (Remote removal of a cap on a small bottle of concentrated permanganate solution at the bottom, or breaking of a small sealed tube containing very concentrated solution, for example.)

**Liquid-liquid.** A large Erlenmeyer flask is prepared about one-half full of dilute acid solution (HCl - 0.1 M). Phenolphthalein indicator ( a few mL) is added, and stirred into the acid. The solution is then allowed to settle for several minutes. One drop of concentrated sodium hydroxide solution is gently placed in the center of the acid solution using a dropping pipette. A pink spot should be created which spreads outwards. An estimate is made of how fast this color spreads, and the experiment repeated until all the solution has turned uniform pink. Distances traveled per unit time can be estimated by placing the flask on a flat ruler with clear mm markings. Discussion should include noting observations that indicate that the solution was not left to settle long enough (uneven spreading), or that it is still subject to some external motion (hand movement, etc.) during the experiment.

*Extension.* A gas-liquid version can be made following the familiar “Voice Activated Chemical Reaction” demonstration including in most books of chemical demonstrations. In this the flask (as above) is half-filled with a solution of bromocresol green indicator in alcohol or alcohol/water. Breathing into the space at the top of the flask (being careful not to disturb the surface) adds carbon dioxide at the surface of the liquid. This will diffuse into the liquid changing the color of the indicator. This can be timed, and a similar discussion of variables should ensue.

**Solid-solid.** Students place a layer of brightly colored sand on the bottom of a clean, clear container with a tight fitting lid (1L or 2L plastic soda bottle will work well). Almost fill the remainder of the bottle with regular sand. Try to find two sands that have very similar grain sizes (unfortunately these materials cannot be reused unless someone invents an easy way to re-separate the colored sand!).

Additionally, students place a thin layer of a second colored sand on the top, leaving some air space. They cap the bottle tightly. Challenge students to estimate how many inversions (or how long in time) it will take to completely mix the sand. Have students who you wish to tire, carry out the inversions, while someone else counts and times. [Note: it may be a good idea to start the class with this demonstration, then it can continue while the others are being carried out.]

In discussion of how long it takes to completely “homogenize” the sands to the class satisfaction, two points can be emphasized. One is that this is only a simulation of formation of a solid solution. In real solutions, there will not be air spaces (though there will be vacancies), and molecules may not be as irregular as sand grains. However, there may be “stickier” forces between the molecules, inhibiting their movement.

The second is that this demonstrates a familiar problem to analytical chemists: how can one be sure that a material sampled is homogenous and thoroughly mixed, if simple mixing as demonstrated here, is

so inefficient. Students may be able to think of ways of speeding the process. Using a larger container (more vacant space) will help considerably. Standard analytical practice, even for solid-liquid or liquid-liquid systems, is to pour the contents into another container, and pour back and forth several times. This usually results in good mixing, better than just stirring with a rod, which may be the first suggestion made by students.

### Background:

Diffusion is an important non-equilibrium phenomenon occurring in solutions. It is the process by which all true solutions are formed from pure substances since there can be no truly instantaneous mixing. Most students should be familiar with what happens if you do not stir a drink such as hot or iced tea after adding sugar crystals, for example, but they may confuse this with the dissolving process itself. The two stage process of making a truly homogenous sugar solution involves the dissolving of the sugar molecules from the surface of the crystal to form water bound clusters, followed by movement of these clusters around and away from the initial solution site as the solution mixes. Students also usually recognize that smells transmitted into the air take some time to reach an individual some distance away. The demonstrations suggested should help to reinforce these notions.

It is also important to develop the notion, however, that just because equilibrium has been reached and no *net* diffusion is occurring, as when the tea all has the same sweetness throughout, or a colored solution is thoroughly mixed, this does not mean that all motion has ceased. It merely means that as many solute molecules are diffusing into a certain volume of solution during each time period as are diffusing out. In other words, representing this as symbols:  $(dX/dt)_{IN} = (dX/dt)_{OUT}$  where  $dX$  represents the change in the number of  $X$  species present *during* the change in time  $dt$ .

In the biological science applications, the phenomenon of diffusion is sometimes confused with osmosis which is related, but not the same. Diffusion is observable when molecules of a solute are not evenly distributed among those of a solvent. There is no semipermeable membrane effecting a differential passage of *selected* molecules as in osmosis. Both phenomena are, however, driven by concentration gradients. The larger the difference in concentration per unit distance, the faster will be the process. The *net* processes cease when there is no longer any concentration gradient. For instance if an uncured pickle is immersed in a poorly stirred salt solution, diffusion controls the concentration of salt solution at the surface of the pickle cells. It also controls the degree to which the cell fluid content is homogenous, but osmosis controls the rate at which water molecules pass through the cell wall/membrane as concentrations of sodium and chloride ions inside and outside each cell begin to equalize.

#### Answers to Student Questions.

1. Particles are moving at the same speed (at a fixed  $T$ ) but, they make for less collisions in a gas than in a liquid or solid, so they *appear* to be moving much faster.

For example, compare a student running across an empty field with one who runs running at the same speed across a crowded field. Which student gets to the other side faster?

2. Changing  $T$  has the largest effect, changing the solvent the next, and changing the solute makes the least difference. (See background information for more details.)

**Variations:**

The teacher prepares and carries out the examples as demonstrations. The students take notes, and then discuss the models as a class discussion. Students should be responsible for making notes on their measurements and observations.

It may be possible to carry out a convincing demonstration using marbles instead of sand. This has the advantage that the marbles can be readily separated and reused, but it seems more remote from the invisible particles we are trying to have the students conceive as moving in real solutions.

Adapted from:  
None.

Item	Consumables	
	Quantity (per lab group)	Activity
benzoic acid (122)	a few grams*	1
benzophenone (49)	a few grams*	1
calcium chloride	a few grams*	2
calcium magnesium acetate (CMA)	a few grams*	2
durene (79)	a few grams*	1
butane lighter	1	1
dilute acid solution (or white vinegar)	50 mL	4
dilute sodium hydroxide solution (0.1 to 1 M)	—	4
dropper bottle with phenolphthalein reagent	1	4
dropper bottle, small, containing a few drops of inexpensive perfume, ammonia solution, or other compound with strong—but not unpleasant—odor	1	4
ethylene glycol	—	3
glutamic acid (98)	a few grams*	1
hydrocinnamic acid (48)	a few grams*	1
hydroquinone(173)	a few grams*	1
ice	—	2
lauric (dodecanoic) acid (44)	a few grams*	1
magnesium chloride	a few grams*	2
maleic acid (140)	a few grams*	1
malonic acid (136)	a few grams*	1
naphthalene (80)	a few grams*	1
phenyl salicylate (salol) (43)	a few grams*	1
potassium acetate	a few grams*	2
potassium permanganate (or crystals of brightly colored, water-soluble chemical)	—	4
propylene glycol	—	3
resorcinol (110)	a few grams*	1
salt (NaCl)	—	2
sand, brightly colored (one or more colors)	25 mL	4
sand, beach (same grain size as colored sand)	500 mL to 1 L	4
sodium chloride	a few grams*	2
stearic acid (70)	a few grams*	1
water, distilled or deionized	—	2, 3, 4

(continued)

<b>Nonconsumables</b>		
<b>Item</b>	<b>Quantity (per lab group)</b>	<b>Activity</b>
boiling point setup	1	3
capillary melting-point tube	1	3
capillary melting-point tubes	2	1
capillary melting-point tubes	2–4	1, 2
chemical balance	1	3
chemical balance, weighs to 0.10 g	1	2
dropper pipette	1	4
dropping pipettes, fine-point	2	3
dropping pipette w/small capillary end	1	2
Erlenmeyer flask	1	4
hand lens	1	2
measuring cylinder (or pipette), 10-mL	1	2
melting point setup	1	1, 2
microscope slides (or small watch glass)	2	1, 2
spatula (or firm drink stirrer)	1	1, 2
test tube, closed end, about 5 mm internal dia.	1	3
test tube, large, (graduated cylinder or 1 L plastic bottle)	1	4

\* indicates amount per class

**Key to activities:**

1. Is It Hot Enough to Melt?
2. Melting Ice is a Good Idea!
3. Boil This!
4. Diffusion in Every Phase

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**Student Readings**

Mergenmeier, Andrew, "New Strategies Can Improve Winter Road Maintenance Operations," *Public Roads*, Washington, D. C.: U. S. Dept. of Transportation, Federal Highway Admin., TD 2.19: 58/4, Spring 1995, pp. 16–17.