CHAPTER 1 | Particles of Matter: Measurement and the Tools of Science

1.1. **Collect and Organize**

In Figure P1.1(a) we are shown "molecules" each consisting of one red sphere and one blue sphere, and in Figure P1.1(b) we have separate blue spheres and red spheres. In each figure we are to identify whether the substance depicted is a solid, liquid, or gas and if the figures show pure elements, compounds, homogeneous mixtures, or heterogeneous mixtures.

Analyze

A pure substance is composed of all the same type of element or compound, not a mixture of two kinds. An element is composed of all the same type of atom, and a compound is composed of two or more types of atoms. In a homogeneous mixture the components are evenly distributed throughout the mixture, giving a uniform appearance to the eye. A heterogeneous mixture contains distinct, observable, individual components. Solids have a definite volume and a highly ordered arrangement where the particles are close together, liquids also have a definite volume but have a disordered arrangement of particles that are close together, and gases have disordered particles that fill the volume of the container and are far apart from each other.

Solve

(a) Because the particles each consist of one red sphere and one blue sphere, all the particles are the same—this is a pure compound. The particles fill the container and are disordered, so these particles are in the gas phase.
(b) Because Figure P1.1(b) shows a mixture of red and blue spheres, this is depicting a mixture of blue-element atoms and red-element atoms. The blue spheres fill the container and are disordered, so these particles are in the gas phase. The red spheres have a definite volume and are slightly disordered, so these particles are in the liquid phase. These two phases are distinct, and we would observe the difference in composition, so this mixture is heterogeneous.

Think About It

Remember that both elements and compounds may be either pure or present in a mixture.

1.2. **Collect and Organize**

In Figure P1.2(a) we are shown "atoms" of only red spheres, and in Figure P1.2(b) we have "molecules" consisting of two red spheres or two blue spheres. In each figure we are to identify whether the substance depicted is a solid, liquid, or gas and if the figures show pure elements, compounds, homogeneous mixtures, or heterogeneous mixtures.

Analyze

A pure substance (whether element or compound) is composed of all the same type of molecule or atom, not a mixture of two kinds. An element is composed of all the same type of atom, and a compound is composed of two or more types of atoms. In a homogeneous mixture the components are evenly distributed throughout the mixture, giving a uniform appearance to the eye. A heterogeneous mixture contains distinct, observable, individual components. Solids have a definite volume and a highly ordered arrangement where the particles are close together, liquids also have a definite volume but have a disordered arrangement of particles that are close together, and gases have disordered particles that fill the volume of the container and are far apart from each other.

Solve

(a) Because all the atoms are of the same type, Figure P1.2(a) depicts a pure element. The particles take up a definite volume and are ordered, so this element is in the solid phase.

(b) Because there is a mixture of blue diatomic molecules and red diatomic molecules, Figure P1.2(b) depicts a mixture of two elements. Both the blue and red diatomic particles fill the volume of the container and are highly disordered; the mixture depicted is in the gas phase, and the mixture is homogeneous.

Think About It

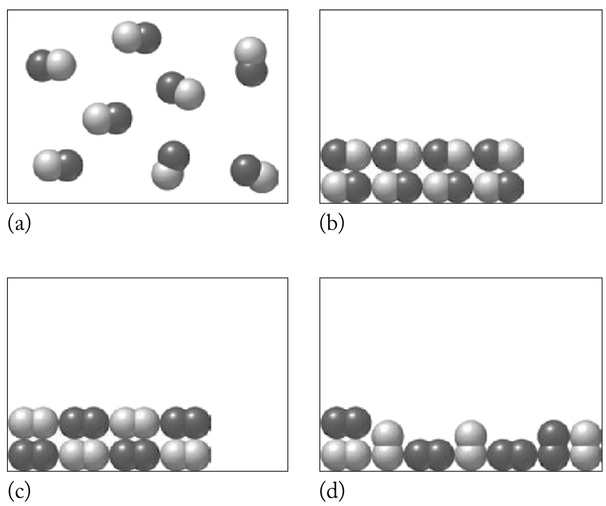
Elements do not need to be present as single atoms. They may be diatomic, as in H_2 or Br_2, or even more highly associated, as in S_8 or P_4.

1.3. **Collect and Organize**

In this question we are to consider whether the reactants as depicted undergo a chemical reaction or a phase change.

Analyze

Chemical reactions involve the breaking and making of bonds in which atoms are combined differently in the products than in the reactants. In considering a possible phase change, solids have a definite volume and a highly ordered arrangement where the particles are close together, liquids also have a definite volume but have a disordered arrangement of particles that are close together, and gases have disordered particles that fill the volume of the container and are far apart from each other. It may help to visualize each scenario, as shown below:



Solve

In Figure P1.3 two pure elements (red–red and blue–blue) in the gas phase recombine to form a compound (red–blue) in the solid phase (ordered array of molecules). Therefore, answer (b) describes the reaction shown.

Think About It

A phase change does not necessarily accompany a chemical reaction. We will learn later that the polarity of the product will determine whether or not a substance will be in the solid, liquid, or gaseous state at a given temperature.

1.4. **Collect and Organize**

In this question we are to consider whether the reactants as depicted undergo a chemical reaction (either recombination or decomposition) or a phase change.

Analyze
Chemical reactions involve the breaking and making of bonds in which atoms are combined differently in the products than in the reactants. In considering a possible phase change, solids have a definite volume and a highly ordered arrangement where the particles are close together, liquids also have a definite volume but have a disordered arrangement of particles that are close together, and gases have disordered particles that fill the volume of the container and are far apart from each other.

Solve
In Figure P1.4 we see that no recombination of the diatomic molecules occurs. The pure element (red–red) condenses to a slightly disordered phase, while the other element (blue–blue) remains in the gas phase. Therefore, answer (a) describes the reaction pictured.

Think About It
Cooling of air in this fashion to different temperatures separates the components of air.

1.5. **Collect and Organize**

From the space-filling model shown, we are to write the formula for the chemical represented.

Analyze
Based on the Atomic Color Palette (see the inside back cover of the textbook), we see that this model contains one hydrogen atom bonded to a carbon atom that is bonded to an oxygen atom and to an O–H unit.

Solve
$HC(O)OH$ or CH_2O_2

Think About It
This model represents formic acid, and the presence of the C(O)OH unit classifies it as a carboxylic acid.

1.6. **Collect and Organize**

From the ball-and-stick model of isopropanol shown, we are to write the chemical formula.

Analyze
Based on the Atomic Color Palette on the inside back cover of the textbook, we see that this model contains two CH_3 units bonded to a central C atom with both an H atom and OH unit.

Solve
$H_3CCH(OH)CH_3$ or C_3H_8O

Think About It
This model represents isopropanol, and the presence of the O–H unit classifies it as an alcohol. Sometimes isopropanol is called “rubbing alcohol.”

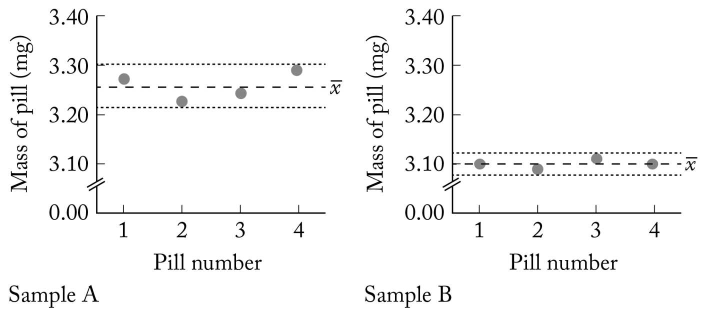
1.7. **Collect and Organize**

This question asks us to consider two sets of data and determine whether each set is precise or accurate.

Analyze
Accuracy described how close the measured value is to the true value. In this case, each pill should weigh 3.25 mg, so the masses must be near this value to be accurate. Precision describes how consistent repeated measurements are with one another.

Solve

Sample A is both accurate and precise. The average of the four masses is near 3.25 mg, and the values appear fairly close to one another. Sample B is precise but not accurate. While the data points are quite similar to one another, the average value is closer to 3.10 mg than 3.25 mg.



Think About It

Accuracy and precision can be somewhat difficult to evaluate without a statistical analysis of the data. In Section 1.10, we discuss how the mean, standard deviation, and confidence intervals can be used to determine if a true value is near our actual data and how similar data points are to one another.

1.8. **Collect and Organize**

In this question, we are asked to examine nine depictions of compounds, elements, and chemical reactions and answer questions about each.

Analyze

A and H are ball-and-stick models, while B, D, and I are space-filling models, and F is a structural formula. C, E, and G are equations with included depictions of the transformation in the real world. Bonds are broken or formed as a chemical reaction proceeds, while a physical process involves a change of phase but no breaking or formation of chemical bonds. Compounds are made from elements and have different types of atoms in them. Elements are composed of atoms that are all the same.

Solve

(a) C is the only chemical reaction depicted. In this reaction, hydrogen peroxide decomposes to form water and oxygen gas.

(b) After consulting the Atomic Color Palette on the inside back cover of the textbook, we see that B contains a central N atom bound to three H atoms, with the molecular formula NH_3. The same molecular formula is depicted in F.

(c) E and G are depictions of physical processes. E depicts the vaporization of N_2, the transformation from the liquid to the gas phase. G depicts the sublimation of I_2, the transformation from the solid to the gas phase.

(d) The pure elements depicted are A (P_4, phosphorus), E (N_2, nitrogen), G (I_2, iodine) and I (O_2, oxygen).

(e) The ball-and-stick depiction in H contains three colored spheres, corresponding to hydrogen (white), carbon (black), and oxygen (red). The molecular formula for this compound is HC(O)H, or CH_2O.

(f) From the molecular formula provided in D (C_5H_{12}), we can tell that this molecule contains 17 atoms, more than any other molecule depicted.

Think About It

C_5H_{12} depicted in D is pentane, a hydrocarbon. The term hydrocarbon is used to describe molecules containing only the elements hydrogen and carbon.

1.9. **Collect and Organize**

In this question, we consider how elements and compounds compare.

Analyze

Compounds are made from elements and have different types of atoms in them. Elements are composed of atoms that are all the same.

Solve

Compounds are different from elements in that they are made up of two or more elements, these elements can be separated from each other (but elements cannot be separated further), and compounds have chemical and physical properties different from the elements that compose them. Compounds are commonly found in nature, but elements rarely are. Compounds are similar to elements in that they are composed of atoms, have definite physical and chemical properties, and can be isolated in pure form.

Think About It

By combining the different elements with each other, we can arrive at many, many compounds that are used as fuels, medicines, plastics, and so on.

1.10. **Collect and Organize**

For this question we are asked to consider what is contained in the space between gas particles.

Analyze

Gases contain individual particles separated by large distances, relative to the size of the particle. All of the mass of the gas is contained in the particles, so no matter can exist in the space between these particles.

Solve

The space between particles is a gas that contains no matter, so there is only empty space present.

Think About It

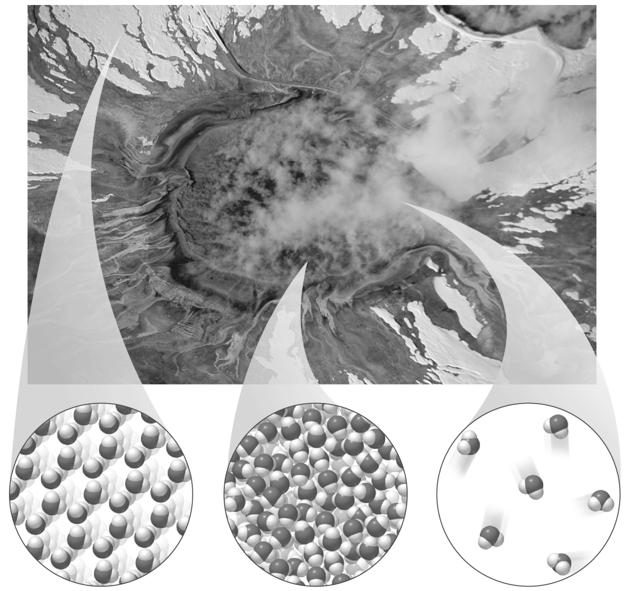
If a sample of gas is compressed to a smaller volume, its mass should remain constant. This supports our claim that the space between gas particles has no mass, and so contains no matter.

1.11. **Collect and Organize**

This question asks us to determine in which phase of matter particles are able to move the most and in which phase of matter particles are able to move the least.

Analyze

In general, the closer particles are to one another, the stronger the forces holding them together. Particles that are held together more tightly move less. A visual inspection of Figure 1.12 allows us to estimate the forces holding water particles together.



(a) Solid (b) Liquid (c) Gas

Solve

Particles are most free to move about in the gas phase. Particles in the solid phase have strong attractive forces between particles, making this the phase in which particles have the least motion.

Think About It

Although liquid particles can be quite close to one another, they are not locked into a regular array (in many solids this is called a crystalline lattice). Liquids are fluid (they can be poured) because individual particles may move past one another.

1.12. **Collect and Organize**

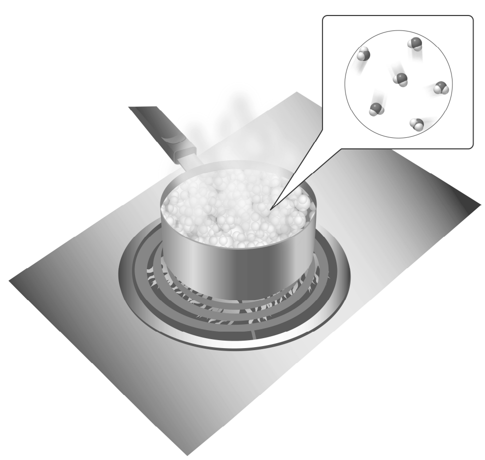
We are to identify the gas contained within bubbles in a pot of boiling water.

Analyze

Few or no bubbles are present in a pot of water before heat is applied. This is a clue that the bubbles are formed from the water itself. When heated sufficiently, liquid water evaporates to form steam, the gaseous form of water.

Solve

The bubbles in a pot of boiling water are water vapor.



Think About It
Though small amounts of gases such as oxygen and nitrogen may be present in tap water, there is not nearly enough to form all the bubbles that are observed when water boils.

1.13. **Collect and Organize**
We are asked to suggest how a small quantity of snow could disappear without melting.

Analyze
Snow is crystalline water in the solid phase. Solids may undergo a phase change to form a liquid (called melting) or a gas (called sublimation).

Solve
The snow sublimed to form water vapor.

Think About It
Direct sublimation of water ice to water vapor happens all the time in your freezer. "Freezer-burned" food items have dried out by losing water in the same manner as the snow in this problem.

1.14. **Collect and Organize**
For the substances listed, we are to determine which are homogeneous.

Analyze
Homogeneous mixtures have the same composition throughout.

Solve
The gold wedding ring, sweat, bottled drinking water, and compressed air in a scuba tank (a, b, c, and e) are homogeneous. Human blood (d) is a heterogeneous mixture.

Think About It
A gold wedding ring is made up of an alloy (a solid solution of one metal dissolved in another) of gold with another metal such as palladium or copper to give the soft gold metal strength and make it less expensive than 100% gold. Human blood is a mixture containing plasma, platelets, and red blood cells, among other components.

1.15. **Collect and Organize**
From the list of properties of sodium, we are to determine which are physical and which are chemical properties.

Analyze
Physical properties are those that can be observed without transforming the substance into another substance. Chemical properties are only observed when one substance reacts with another and therefore is transformed into another substance.

Solve
Density, melting point, thermal and electrical conductivity, and softness (a–d) are all physical properties, whereas tarnishing and reaction with water (e, f) are both chemical properties.

Think About It
Because the density of sodium is less than that of water, a piece of sodium will float on water as it reacts.

1.16. **Collect and Organize**

From the list of properties of hydrogen gas, we are to determine which are physical and which are chemical properties.

Analyze

Physical properties are those that can be observed without transforming the substance into another substance. Chemical properties are only observed when one substance reacts with another and therefore is transformed into another substance.

Solve

Density, boiling point, and electrical conductivity (a, c, d) are all physical properties, whereas the reaction of hydrogen with oxygen (b) is a chemical property.

Think About It

Because the density of hydrogen gas is lower than that of any other gas, a lightweight balloon filled with hydrogen will float in air like the more familiar helium balloon.

1.17. **Collect and Organize**

We are asked to determine which of the given species is *not* a pure substance.

Analyze

All matter may be classified as either a pure substance or a mixture. If a substance may be separated by physical means, it is not a pure substance.

Solve

Air may be separated into simpler components, so it cannot be a pure substance. Nitrogen gas, oxygen gas, and argon gas are all pure elements. Pure sodium chloride contains only NaCl.

Think About It

Air is a mixture of nitrogen gas, oxygen gas, argon gas, and other pure substances.

1.18. **Collect and Organize**

We are asked to determine which of the given forms of matter is a pure substance.

Analyze

Pure substances cannot be separated by a physical process.

Solve

Sweat, blood, brass, and milk are mixtures of more than one substance. Only sucrose (table sugar) is a pure compound that cannot be physically separated into simpler components.

Think About It

Sweat, blood, and milk all contain mostly water; the other components are what make them different from one another. Brass is an alloy, meaning it is a mixture of two solids. Sucrose (table sugar) can only be separated into other components by chemical means, such as respiration and combustion.

1.19. **Collect and Organize**

From the given list, we are asked to identify the elements.

Analyze

An element is a pure substance that cannot be broken down into simpler components by any process.

Solve

Only Cl_2 is a pure element. The other species listed are compounds of more than one type of element.

Think About It

The elements are listed in the periodic table of the elements. Some elements, such as chlorine, form diatomic molecules (like Cl_2) in their pure form.

1.20. **Collect and Organize**

From the given list, we are asked to identify the species that are *not* elements.

Analyze

An element is a pure substance that cannot be broken down into simpler components by any process. Any pure substance that is composed of multiple elements is a compound.

Solve

Iodine (I_2) and sulfur (S) are pure elements. ClF is the only substance that is composed of more than one type of element; it is classified as a compound, not an element. Ozone (O_3) contains only one type of element, but it is not itself a pure element. Pure oxygen gas in its elemental state is a diatomic molecule, O_2.

Think About It

Despite containing multiple atoms in their pure elemental state, I_2 and S_8 are still elements because they consist of only one type of atom.

1.21. **Collect and Organize**

We are asked to determine which of the given mixtures is homogeneous.

Analyze

A homogeneous mixture has no visible boundaries and contains a uniform distribution of components. We should picture each of these items and decide if their components are uniformly distributed.

Solve

Only filtered water is uniformly distributed, and so it is the only homogeneous mixture.

Think About It

All of the other species listed are heterogeneous mixtures since they have components that are visually different or are not uniformly distributed. Clouds may take many shapes and densities, so they do not exhibit a uniform distribution.

1.22. **Collect and Organize**

From the list below, we are asked to identify the species that are heterogeneous mixtures.

Analyze

In a heterogeneous mixture, visible boundaries may be observed, and components are not uniformly distributed. We should first decide if these items are mixtures or pure substances, then classify the mixtures as homogeneous or heterogeneous.

Solve
Muddy river water is the only heterogeneous mixture from this list. While the mixture may appear uniform, suspended mud particles stick together in an irregular fashion.

Think About It
Air, sugar dissolved in water, and brass are homogeneous mixtures because all components are spread uniformly throughout the samples. Suspended mud particles will settle from a sample of muddy water if the sample is allowed to sit undisturbed for long enough. Table salt is a pure substance consisting only of Na and Cl.

1.23. **Collect and Organize**
We are to determine which mixtures can be separated into their components by filtering.

Analyze
Filtration is used to separate suspended solids from a liquid or gas; this technique will only work if our sample is a heterogeneous mixture containing a solid suspended in a liquid or a gas.

Solve
Sand and water may be separated using filtration. The sand will be trapped on the filter, and the water will pass through.

Think About It
The other mixtures listed in the question are homogeneous and cannot be separated by filtration.

1.24. **Collect and Organize**
In this question we consider whether filtration would be a suitable method to separate a protein from the other components of blood plasma.

Analyze
Filtration separates suspended solids from a heterogeneous mixture containing a liquid. Blood plasma is a solution (homogeneous mixture) containing water and dissolved proteins.

Solve
Filtration will not remove the proteins in blood plasma as individual proteins are too small. We are told that these proteins are dissolved, not suspended.

Think About It
Filtration is not appropriate because the blood plasma is a homogeneous solution.

1.25. **Collect and Organize**
From the list of properties of formaldehyde, we are asked to determine which one is a chemical property.

Analyze
Chemical properties are observed when a chemical reaction takes place. We should identify which property relates to a chemical reaction between formaldehyde and another substance.

Solve
Formaldehyde must react with the air in order to burn. Flammability (c) is a chemical property.

Think About It
Smell, solubility, phase of matter, and color are all physical properties. We can observe these properties without destroying (reacting) the formaldehyde.

1.26. **Collect and Organize**
From the list of properties of silver, we are asked to determine which one is a physical property.

Analyze
Physical properties can be observed without transforming one substance into another.

Solve
We may observe that silver will sink in water without converting it into another substance, making this (d) a physical property.

Think About It
Tarnish forms on silver when it reacts with compounds in the air around it. Tarnishing or other reactions of silver are chemical events.

1.27. **Collect and Organize**
We are to explain if an extensive property can be used to identify a substance.

Analyze
An extensive property is one that, like mass, length, and volume, is determined by size or amount.

Solve
Extensive properties will change with the size of the sample and therefore cannot be used to identify a substance.

Think About It
We could, for example, have the same mass of feathers and lead, but their mass alone will not tell us which mass measurement belongs to which—the feathers or the lead.

1.28. **Collect and Organize**
Of the properties listed, we are to choose which are intensive properties and which are extensive properties.

Analyze
An intensive property is not dependent on the size or amount of the sample, while an extensive property does depend on the size or amount of a sample.

Solve
Of the properties listed, density, the temperature at which a phase change will occur, and the mass of a single molecule (a, b, and d) are all intensive properties. The mass of water in your body (c) is an extensive property because the value depends on the size of your body. The rate at which water flows over Niagara Falls (e) is also an extensive property because the value will vary depending on recent rainfall and even the season.

Think About It
Intensive properties may be used to identify a substance as they will not change based on the size or quantity of the sample used. Intensive properties are related to chemical interactions between atoms and molecules in the substance that will not change from sample to sample.

1.29. **Collect and Organize**

In this question we think about the information needed to formulate a hypothesis.

Analyze

A hypothesis is a tentative explanation for an observation.

Solve

To form a hypothesis we need at least one observation, experiment, or idea (from examining nature).

Think About It

A hypothesis that is tested and shown to be valid can become a theory.

1.30. **Collect and Organize and Analyze**

In this question we consider how a hypothesis becomes a theory.

Solve

A theory is formed from a hypothesis when the hypothesis has been extensively tested by many observations and experiments. A theory is the best (current) possible explanation that is extensively supported by experimentation.

Think About It

A theory, tested over time, may be elevated further to become a scientific law.

1.31. **Collect and Organize**

We are to consider whether we can disprove a hypothesis.

Analyze

A hypothesis is a tentative explanation for an observation.

Solve

It is possible to disprove a scientific hypothesis. In fact, many experiments are designed to do just that as the best test of the hypothesis's validity.

Think About It

It is even possible to disprove a theory (albeit harder to do so) or cause a theory to be modified when new evidence, a new experimental technique, or new data from a new instrument give observations that are counter to the explanation stated by the theory.

1.32. **Collect and Organize**

In this question we are asked why the idea that all matter consists of atoms was described as a philosophy in ancient Greece, but has been referred to as the atomic theory since the 1800s.

Analyze

The ancient Greeks first proposed that matter was composed of small units called atoms. The term philosophy refers to a set of beliefs, while in science, a theory describes a proposal that has been validated by many experiments.

Solve

The key distinction is that the ancient Greeks were not able to demonstrate that their proposal was correct by conducting scientific experiments. Today, the atomic theory is universally accepted because of the sheer volume of evidence from various experiments and observations that support the hypothesis that matter is composed of atoms.

Think About It
In the 1980s, the techniques of scanning tunneling microscopy and atomic force microscopy were developed to give us indirect "pictures" of atoms and molecules.

1.33. **Collect and Organize**
We are to define *theory* as used in conversation.

Analyze
Theory in everyday conversation has a different meaning than it does in science.

Solve
Theory in normal conversation is someone's idea or opinion or speculation that can easily be changed and may not have much evidence or many arguments to support it.

Think About It
A *theory* in science is a generally accepted and highly tested explanation of observed facts.

1.34. **Collect and Organize**
We consider whether a theory can be proven.

Analyze
Theory in science is the best (current) possible explanation that is extensively supported by experimentation and observations.

Solve
Theory is nearly equivalent to fact in science, without being the absolute truth. A theory is hard to prove absolutely but has many, many supporting experiments whose observations strongly support the theory.

Think About It
The results of one experiment that counter the explanation for a phenomenon explained by a theory could disprove a theory, so theories may be toppled and replaced with new explanations and theories.

1.35. **Collect and Organize**
We are to compare SI units to English units.

Analyze
SI units are based on a decimal system to describe basic units of mass, length, temperature, energy, and so on, whereas English units vary.

Solve
SI units, which were based on the original metric system, can be easily converted into a larger or smaller unit by multiplying or dividing by multiples of 10. English units are more complicated to manipulate. For example, to convert miles to feet you have to know that there are 5280 feet in 1 mile, and to convert gallons to quarts you have to know that 4 quarts are in 1 gallon.

Think About It
Once you can visualize a meter, a gram, and a liter, using the SI system is quite convenient.

1.36. **Collect and Organize and Analyze**

In this question we are to suggest two reasons why SI units are not more widely used in the United States.

Solve

English units instead of SI units are used everywhere in the United States because many of our manufacturing facilities have been built to make parts in inches or to bottle liquids in gallons. It has also been difficult for people used to buying and measuring in the English units to convert their thinking so as to visualize a kilometer instead of a mile or a liter instead of a quart.

Think About It

The only widespread everyday use of an SI unit in the United States is the 2 L soda bottle.

1.37. **Collect and Organize**

This question asks us to convert a mass in pounds to grams. There are approximately 2 pounds in a kilogram, so the mass should be just under 1000 grams.

Analyze

We may convert this mass using the following conversion:

$$\frac{453.4 \text{ g}}{1 \text{ lb}}$$

Solve

$$1.65 \text{ lb} \times \frac{453.4 \text{ g}}{1 \text{ lb}} = 748 \text{ g} \quad \text{or} \quad 7.48 \times 10^2 \text{ g}$$

Think About It

A gram is a relatively small unit of mass, so it makes sense that 1.65 pounds would contain 748 grams (just under one kilogram).

1.38. **Collect and Organize**

This question asks us to convert a mass in grams to pounds. There are approximately 2 pounds in a kilogram, so the mass should be just under 2 pounds.

Analyze

We can convert this mass using the following conversion factor:

$$\frac{453.4 \text{ g}}{1 \text{ lb}}$$

Solve

$$765.4 \text{ g} \times \frac{1 \text{ lb}}{453.4 \text{ g}} = 1.688 \text{ lb}$$

Think About It

Since the gram is a relatively small unit of mass, a relatively large number of grams should only amount to a few (just under 2) pounds.

1.39. **Collect and Organize**

This question asks us to convert a volume in gallons to milliliters. Remember that the prefix *milli-* can be read as 10^{-3}. There are approximately 4 liters in a gallon, so our value (just under 2.5 gallons) should be approximately 10 liters, or 10,000 milliliters.

Analyze

We can convert gallons to liters, then liters to milliliters, using the following conversions:

$$\frac{3.7854\ \text{L}}{1\ \text{gal}},\ \frac{1000\ \text{mL}}{1\ \text{L}}$$

Solve

$$2.44\ \text{gal} \times \frac{3.7854\ \text{L}}{1\ \text{gal}} \times \frac{1000\ \text{mL}}{1\ \text{L}} = 9.24 \times 10^{3}\ \text{mL}$$

Think About It

Since the milliliter is a relatively small unit of volume, it makes sense that a large unit like a gallon would contain many milliliters. This volume is just under our estimate of 10 liters. You may also have arrived at the same answer using the conversion

$$\frac{1\ \text{mL}}{1 \times 10^{-3}\ \text{L}}$$

1.40. **Collect and Organize**

This question asks us to convert a volume in milliliters to gallons. Remember that the prefix *milli-* can be read as 10^{-3}. There are approximately 4 liters in a gallon, so our value should be much less than 1 gallon.

Analyze

We can convert milliliters to liters, then liters to gallons, using the following conversions:

$$\frac{1\ \text{gal}}{3.7854\ \text{L}},\ \frac{1\ \text{L}}{1000\ \text{mL}}$$

Solve

$$108\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} \times \frac{1\ \text{gal}}{3.7854\ \text{L}} = 0.0285\ \text{gal} = 2.85 \times 10^{-2}\ \text{gal}$$

Think About It

108 mL is approximately 4 fluid ounces, which is significantly smaller than a gallon.

1.41. **Collect and Organize**

We are asked to compare two lengths, expressed using different units. We should express these values using the same units to find the larger number.

Analyze

We can convert feet to inches, and then add this value to the 11 remaining inches to determine Peter's height in inches. We may then convert inches to centimeters and compare Peter and Paul's heights on the same scale. The conversions we will use are

$$\frac{12\ \text{in}}{1\ \text{ft}},\ \frac{2.54\ \text{cm}}{1\ \text{in}}$$

Solve

First, we will convert Peter's height to inches:

$$\left(5\text{ ft}\times\frac{12\text{ in}}{1\text{ ft}}\right)+11\text{ in}=71\text{ in}$$

Peter's height in centimeters is:

$$71\text{ in}\times\frac{2.54\text{ cm}}{1\text{ in}}=180\text{ cm}=1.8\times10^{2}\text{ cm}$$

Peter is slightly taller than Paul.

Think About It

Because their heights are so similar, we can only tell who is taller because of the number of significant figures in Paul's height. If we had measured Paul's height with only two significant figures, the two would not appear different. You might also convert Paul's height into inches and find that it is 69.3 in, which is also slightly smaller than Peter's height.

1.42. **Collect and Organize**

We are asked to compare two depths, expressed in different units. After expressing these in the same units, the larger value will correspond to the deep end of the pool.

Analyze

We can convert inches to centimeters, then centimeters to meters using the following conversions:

$$\frac{2.54\text{ cm}}{1\text{ in}},\ \frac{1\text{ m}}{100\text{ cm}}$$

Solve

In meters, the depth of the end with a depth of 72 inches is

$$72\text{ in}\times\frac{2.54\text{ cm}}{1\text{ in}}\times\frac{1\text{ m}}{100\text{ cm}}=1.8\text{ m}$$

B is the deep end.

Think About It

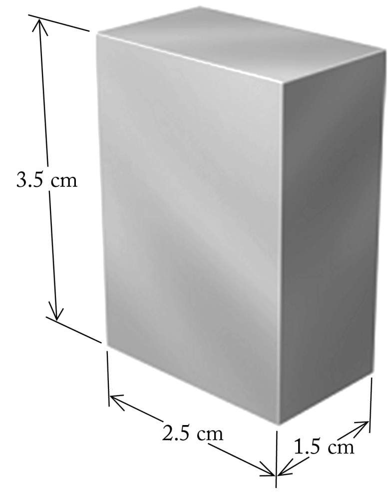
You could also convert the depth at end A to inches and compare the two values. End A is 44 inches deep.

1.43. **Collect and Organize**

In this problem we need to use the density of magnesium to find the mass of a specific size block of the metal.

Analyze

Density is defined as the mass of a substance per unit volume. The density of magnesium is given in Appendix 3 as 1.738 g/cm^3. We have to find the volume of the block of magnesium by multiplying the length by the height by the depth (the value will be in cm^3).



We can then find mass using the following formula:

$$\text{Mass (g)} = \text{density (g/cm}^3) \times \text{volume (cm}^3)$$

Solve

The volume of the block of magnesium is:

$$2.5\text{ cm} \times 3.5\text{ cm} \times 1.5\text{ cm} = 13\text{ cm}^3$$

Therefore, the mass of the block is

$$13\text{ cm}^3 \times \frac{1.738\text{ g}}{1\text{ cm}^3} = 23\text{ g}$$

Think About It

The mass of a sample depends on how much there is of a substance. In this case, we have about 23 grams. As a quick estimate, a block of magnesium of about 10 cm^3 would weigh 10 times that of 1 cm^3, or about 17 grams. Because we have more than 10 cm^3 of this sample and the density is a little greater than 1.7 g/cm^3, our answer of 23 grams is reasonable.

1.44. **Collect and Organize**

In this problem we need to use the density of osmium to find the mass of a specific size block of the metal. To find out whether we could lift it with one hand, we may also have to convert it into pounds since that is the unit we are more closely familiar with in the United States.

Analyze

Density is defined as the mass of a substance per unit volume. In Appendix 3 the density of osmium is given as 22.57 g/cm^3. We have to find the volume of the block of osmium by multiplying the length by the height by the depth (the value will be in cm^3). We can then find mass through the following formula:

$$\text{Mass (g)} = \text{density (g/cm}^3) \times \text{volume (cm}^3)$$

We can use the conversion of grams to pounds for the comparison:

$$\frac{453.6\text{ g}}{1\text{ lb}}$$

Solve

The volume of the block of osmium is

$$6.5\ \text{cm} \times 9.0\ \text{cm} \times 3.25\ \text{cm} = 190\ \text{cm}^3$$

Therefore, the mass of the block is

$$190\ \text{cm}^3 \times \frac{22.57\ \text{g}}{1\ \text{cm}^3} = 4300\ \text{g}$$

To convert this into the more familiar pounds:

$$4300\ \text{g} \times \frac{1\ \text{lb}}{453.6\ \text{g}} = 9.5\ \text{lb}$$

Think About It

Nearly 10 pounds is fairly heavy, but it could be lifted with one hand. The block of osmium, though, will be surprisingly heavy because its volume is relatively small.

1.45. **Collect and Organize**

This problem asks us to determine the conversion factors to convert the given units.

Analyze

The conversions that we need include picoseconds and femtoseconds to seconds, milligrams and kilograms to grams, and centimeters to meters.

$$\frac{10^{-12}\ \text{s}}{1\ \text{ps}},\ \frac{10^{-15}\ \text{s}}{1\ \text{fs}},\ \frac{1\ \text{kg}}{1000\ \text{g}},\ \frac{1\ \text{g}}{1000\ \text{mg}},\ \frac{1\ \text{m}}{100\ \text{cm}}$$

The density of titanium is given in Appendix 3 as 4.54 g/cm^3.

Solve

(a)

$$1\ \text{ps} \times \frac{10^{-12}\ \text{s}}{1\ \text{ps}} \times \frac{1\ \text{fs}}{10^{-15}\ \text{s}} = 1 \times 10^3\ \text{fs}$$

As a conversion factor: $\frac{1\ \text{ps}}{1 \times 10^3\ \text{fs}}$

(b)

$$1\ \text{kg} \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \frac{1000\ \text{mg}}{1\ \text{g}} = 1 \times 10^6\ \text{mg}$$

As a conversion factor: $\frac{1\ \text{kg}}{1 \times 10^6\ \text{mg}}$

(c)

$$1\ \text{kg Ti} \times \frac{10^3\ \text{g Ti}}{1\ \text{kg Ti}} \times \frac{1\ \text{cm}^3}{4.54\ \text{g Ti}} \times \left(\frac{1\ \text{m}}{100\ \text{cm}}\right)^3 = 2.20 \times 10^{-4}\ \text{m}^3$$

As a conversion factor: $\frac{1\ \text{kg Ti}}{2.20 \times 10^{-4}\ \text{m}^3}$

Think About It

The density of a substance is an intensive property and will remain constant for that material. Though the values differ, the density of titanium is the same whether it is expressed in g/cm^3 or kg/m^3.

1.46. **Collect and Organize**

This problem asks for a simple conversion of length from meters to miles.

Analyze

The conversions that we need include meters to kilometers and kilometers to miles.

$$\frac{1\ \text{km}}{1000\ \text{m}}, \quad \frac{0.6214\ \text{mi}}{1\ \text{km}}$$

Solve

$$4.0 \times 10^{3}\ \text{m} \times \frac{1\ \text{km}}{1000\ \text{m}} \times \frac{0.6214\ \text{mi}}{1\ \text{km}} = 2.5\ \text{mi}$$

Think About It

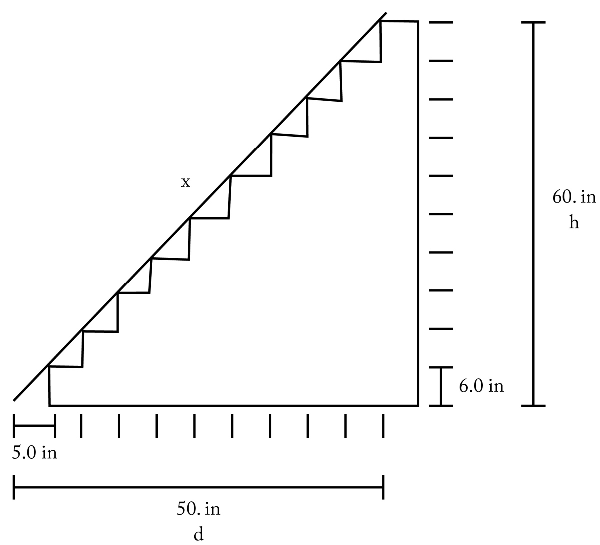
The answer is reasonable because 4000 meters would be a little over 2 miles when estimated. It is surprising, though, for a natural piece of silk to be that long.

1.47. **Collect and Organize**

This problem asks about the total distance along a set of stairs with ten steps.

Analyze

Each step is 6.0 inches higher than the previous step and 5.0 inches deep. A diagram is helpful to visualize this set of stairs.



The conversion from inches to cm is given by

$$\frac{2.54\ \text{cm}}{1\ \text{in}}$$

Solve

The total height of the steps is

$$10\ \text{steps} \times \frac{6.0\ \text{in}}{1\ \text{step}} = 60.\ \text{in}$$

The total depth of the steps is

$$10 \text{ steps} \times \frac{5.0 \text{ in}}{1 \text{ step}} = 50. \text{ in}$$

The distance along the edge of the steps from the bottom to the top is

$$60. \text{ in} + 50. \text{ in} = 110. \text{ in}$$

$$110. \text{ in} \times \frac{2.54 \text{ cm}}{1 \text{ in}} = 279 \text{ cm}$$

Think About It

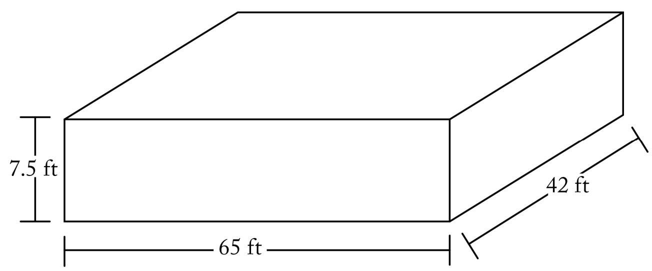
We could also have converted the distance to feet by dividing 110. inches by 12 in/ft. This distance would be 9.17 ft.

1.48. **Collect and Organize**

This problem asks us to determine the volume of a pool in cubic inches.

Analyze

Volume is determined by multiplying the depth, width, and length of the pool.



Since the final units are to be in^3, we may begin by converting the given dimensions in feet to inches. The conversion that we need is

$$\frac{12 \text{ in}}{1 \text{ ft}}$$

Solve

$$7.5 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} = 90 \text{ in}$$

$$42 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} = 504 \text{ in (2 sf)}$$

$$65 \text{ ft} \times \frac{12 \text{ in}}{1 \text{ ft}} = 780 \text{ in (2 sf)}$$

$$\text{Volume} = \text{depth} \times \text{width} \times \text{length}$$

$$= 90 \text{ in} \times 504 \text{ in} \times 780 \text{ in} = 3.5 \times 10^7 \text{ in}^3$$

Think About It

Thirty-five million cubic inches may seem like a large volume of water for a swimming pool, but the units are somewhat small. We may convert this to a volume of approximately 1.5×10^5 gallons, a more common unit for large quantities of liquid.

1.49. **Collect and Organize**

To compute the runner's speed we have to use this definition: speed = change in distance/change in time. In the marathon runner's case we are given distance in miles and time in hours plus additional minutes. The first calculation of speed, therefore, in miles per hour will not require any unit conversion. That result will be used to compute the runner's speed in meters per second using conversions for miles to meters and hours to seconds.

Analyze

The equation to compute speed is given by

$$\text{Speed} = \frac{\Delta \text{ distance}}{\Delta \text{ time}}$$

Because the time is given as 3 hours 34 minutes, we will have to convert the 34 minutes into a part of an hour using the fact that 1 hour = 60 minutes. We can then divide the marathon distance by this time in hours to get the speed in miles per hour.

To convert this speed to meters per second, we can use the following conversions:

$$\frac{1 \text{ km}}{1000 \text{ m}} \text{ and } \frac{0.6214 \text{ mi}}{1 \text{ km}}$$

$$\frac{1 \text{ min}}{60 \text{ s}} \text{ and } \frac{1 \text{ h}}{60 \text{ min}}$$

Solve

First, the number of hours for the runner to complete the marathon is

$$3 \text{ h} + \left(34 \text{ min} \times \frac{1 \text{ h}}{60 \text{ min}}\right) = 3.57 \text{ h}$$

(a) The speed in miles per hour is

$$\text{speed} = \frac{26.2 \text{ mi}}{3.57 \text{ h}} = 7.34 \text{ mi/h}$$

(b) Converting this speed to meters per second gives:

$$7.34 \frac{\text{mi}}{\text{h}} \times \frac{1 \text{ km}}{0.6214 \text{ mi}} \times \frac{1000 \text{ m}}{1 \text{ km}} \times \frac{1 \text{ h}}{60 \text{ min}} \times \frac{1 \text{ min}}{60 \text{ s}} = 3.28 \text{ m/s}$$

Think About It

Both of these values seem reasonable. A walking pace is about 3 miles per hour, so running could be imagined to be twice that fast. Also, 3 meters per second can be run easily by a fast runner.

1.50. **Collect and Organize**

For this problem we need to convert the distance of the Olympic mile (1500 m) from meters to miles and then to feet, compare that distance to a U.S. mile using a ratio, and then convert that into a percentage.

Analyze

For converting the distance we can make use of these conversions:

$$\frac{1\text{ km}}{1000\text{ m}},\ \frac{0.6214\text{ mi}}{1\text{ km}},\ \text{and}\ \frac{5280\text{ ft}}{1\text{ mi}}$$

To determine the percentage the Olympic mile distance is compared to the U.S. mile, we will use

$$\%\ \text{distance} = \frac{\text{Olymoic mile distance in feet}}{5280\text{ ft}} \times 100$$

Solve

$$1500\text{ m} \times \frac{1\text{ km}}{1000\text{ m}} \times \frac{0.6214\text{ mi}}{1\text{ km}} \times \frac{5280\text{ ft}}{1\text{ mi}} = 4921\text{ ft}$$

$$\%\ \text{distance} = \frac{4921\text{ ft}}{5280\text{ ft}} \times 100 = 93.20\%$$

Think About It

This calculation shows that the Olympic mile is just a little bit shorter than the actual mile.

1.51. **Collect and Organize**

To find the Calories burned by the wheelchair marathoner in a race, we can first find the number of hours the race will be for the marathoner at the pace of 13.1 miles per hour. The Calories burned can then be calculated from that value and the rate at which the marathoner burns Calories.

Analyze

The time it takes for the marathoner to complete the race will be given by

$$\text{Time to complete the marathon} = \frac{\text{distance of the marathon}}{\text{pace of the marathoner}}$$

The Calories burned will be computed by

$$\text{Calories burned} = \frac{\text{Calories burned}}{\text{hour}} \times \text{time of the marathon race}$$

Solve

$$\text{Time to complete the marathon} = 26.2\text{ mi} \times \frac{1\text{ h}}{13.1\text{ mi}} = 2.00\text{ h}$$

$$\text{Calories burned} = \frac{665\text{ Cal}}{\text{h}} \times 2.00\text{ h} = 1.33 \times 10^3\text{ Cal}$$

Think About It

This problem could be solved without touching a calculator! Because it takes the marathoner 2.00 h to complete the race, the Calories she burns is simply twice the number of Calories she burns in one hour.

1.52. **Collect and Organize**

This problem involves the conversion of an interstellar distance from light-years to kilometers.

Analyze

The conversions needed are

$$\frac{1 \text{ year}}{365.25 \text{ d}}, \frac{1 \text{ d}}{24 \text{ h}}, \frac{1 \text{ h}}{60 \text{ min}}, \frac{1 \text{ min}}{60 \text{ s}}, \frac{1 \text{ km}}{1000 \text{ m}}$$

From Appendix 2, the speed of light in a vacuum is 2.99792458×10^8 m/s. A light year is the distance light travels in one year's time.

Solve

One light-year represents a distance of

$$1 \text{ year} \times \frac{365.25 \text{ d}}{1 \text{ year}} \times \frac{24 \text{ h}}{1 \text{ d}} \times \frac{60 \text{ min}}{1 \text{ h}} \times \frac{60 \text{ s}}{1 \text{ min}} \times \frac{2.99792458 \times 10^8 \text{ m}}{1 \text{ s}} \times \frac{1 \text{ km}}{1000 \text{ m}} = 9.5 \times 10^{12} \text{ km}$$

In km, the distance to Proxima Centauri is

$$4.3 \text{ light-years} \times \frac{9.5 \times 10^{12} \text{ km}}{1 \text{ light-year}} = 4.1 \times 10^{13} \text{ km}$$

Think About It

Interstellar distances are so large that it can be difficult to visualize relative to everyday objects. The same is true when considering the size of atoms and molecules relative to everyday objects.

1.53. **Collect and Organize**

To answer this question we need to compute the mass of a copper sample that is 125 cm^3 in volume using the density of copper. Next, we use that mass to find out what volume (in cm^3) that mass of gold would occupy.

Analyze

We need the density both of copper and of gold from Appendix 3 to make the conversions from volume to mass (for copper) and then from mass to volume (for gold). These densities are 8.96 g/mL for copper and 19.3 g/mL for gold. One milliliter is equivalent to 1 cm^3, so the densities are 8.96 g/cm^3 and 19.3 g/cm^3, respectively. The density formulas that we need are

$$\text{Mass of copper} = \text{density of copper} \times \text{volume}$$

and

$$\text{Volume of gold} = \frac{\text{mass}}{\text{density of gold}}$$

Solve

$$\text{Mass of copper} = 8.96 \text{ g/cm}^3 \times 125 \text{ cm}^3 = 1120 \text{ g}$$

$$\text{Volume of gold} = \frac{1120 \text{ g}}{19.3 \text{ g/cm}^3} = 58.0 \text{ cm}^3$$

Think About It

Because gold is more than twice as dense as copper, we would expect the volume of a gold sample to have about half the volume of that of the same mass of copper.

1.54. **Collect and Organize**

We have to consider how much air (in grams) the balloon holds when cool. Then, we can use that mass combined with the new larger volume of the heated balloon to find the new density.

Analyze

Density is mass per volume, so we can obtain the mass of air in the cool balloon through

$$\text{Mass} = \text{density of air} \times \text{volume of cool balloon}$$

Then, using this mass we can find the new density through

$$\text{Density} = \frac{\text{mass of air}}{\text{volume of heated balloon}}$$

Solve

$$\text{Mass} = \frac{1.20 \text{ g}}{\text{L}} \times 1.00 \times 10^6 \text{ L} = 1.20 \times 10^6 \text{ g}$$

$$\text{Density of heated ballon} = \frac{1.20 \times 10^6 \text{ g}}{1.09 \times 10^6 \text{ L}} = 1.10 \text{ g/L}$$

Think About It

Because the volume of the heated balloon is larger than that of the cooled balloon, we would expect that the density of the air has decreased upon expanding the balloon.

1.55. **Collect and Organize**

Using the density of mercury, we can find the volume of 1.00 kg of mercury.

Analyze

The density of mercury is given in Appendix 3 as 13.546 g/mL. Because this property is expressed in grams per milliliter, not kilograms per milliliter, we have to convert kilograms into grams using the conversion:

$$\frac{1000 \text{ g}}{1 \text{ kg}}$$

Once we have the mass in grams, we can use the rearranged formula for density to find volume:

$$\text{Volume of mercury (mL)} = \frac{\text{mass of mercury (g)}}{\text{density of mercury (g/mL)}}$$

Solve

$$1.00 \text{ kg} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 1.00 \times 10^3 \text{ g}$$

$$\text{Volume of mercury} = \frac{1.00 \times 10^3 \text{ g}}{13.546 \text{ g/mL}} = 73.8 \text{ mL}$$

Think About It

This is a fairly small amount that weighs one kilogram. This is due to the relatively high density of mercury.

1.56. **Collect and Organize**

For this problem we need to make a comparison of the student's measurement of the density of her piece of jewelry to the known density of silver.

Analyze

The density of the piece of jewelry can be calculated by dividing the mass of the piece of jewelry by its volume according to the formula for density.

Solve

$$\text{Density of the piece of jewelry} = \frac{3.17\ \text{g}}{0.3\ \text{mL}} = 10.\ \text{g/mL}$$

The density of silver is 10.50 g/mL. Because these densities match very closely, the jewelry could be made of pure silver.

Think About It

The difference in the densities between the piece of jewelry and pure silver is less than 5%. A better comparison of densities would require more significant digits for the measurements of mass and volume of the piece of jewelry.

1.57. **Collect and Organize**

Given the average density of human blood, we are asked to determine the mass of liquid occupying a volume of 5.5 L.

Analyze

Given the density in g/mL, we will need to convert the volume from L to mL so that our units cancel. We will also need to convert the mass in grams to a mass in ounces. The conversion factors we will need are

$$\frac{1000\ \text{mL}}{1\ \text{L}},\ \frac{1\ \text{oz}}{28.35\ \text{g}}$$

Solve

(a)
$$5.5\ \text{L} \times \frac{1000\ \text{mL}}{1\ \text{L}} \times \frac{1.06\ \text{g}}{1\ \text{mL}} = 5.8 \times 10^3\ \text{g}$$

(b)
$$5.5\ \text{L} \times \frac{1000\ \text{mL}}{1\ \text{L}} \times \frac{1.06\ \text{g}}{1\ \text{mL}} \times \frac{1\ \text{oz}}{28.35\ \text{g}} = 2.1 \times 10^2\ \text{g}$$

Think About It

It makes sense that this mass would have a smaller value when expressed in ounces rather than grams. Ounces are a larger unit of mass than grams. One pound contains 16 ounces and approximately 454 grams.

1.58. **Collect and Organize**

In this problem we are asked to calculate the Earth's density in grams per cubic centimeter given the mass in kilograms (we have to convert to grams) and the volume of the Earth in cubic kilometers (we have to convert to cubic centimeters).

Analyze

We may convert the units for mass and volume using the following conversion factors:

$$\frac{1000\ \text{g}}{1\ \text{kg}},\ \frac{1000\ \text{m}}{1\ \text{km}},\ \frac{100\ \text{cm}}{1\ \text{m}}$$

It is important to remember that even though there are 100 cm in 1 m, there are $(100\ \text{cm})^3$ or $1 \times 10^6\ \text{cm}^3$ in $1\ \text{m}^3$. Both the value and the unit must be cubed to correctly solve this problem. Density is equal to the mass of the object, divided by the volume that object occupies.

Solve

$$d = \frac{m}{V} = \frac{5.98 \times 10^{24}\ \text{kg}}{1.08 \times 10^{12}\ \text{km}^3} \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \left(\frac{1\ \text{km}}{1000\ \text{m}}\right)^3 \times \left(\frac{1\ \text{m}}{100\ \text{cm}}\right)^3 = 5.54\ \text{g/cm}^3$$

Think About It

Even though the Earth's surface is mostly covered by water (with a density of approximately $1\ \text{g/cm}^3$), the Earth contains a large quantity of heavier matter such as iron and other metals. A density of $5.54\ \text{g/cm}^3$ is a reasonable value since it is not lighter than water and not hundreds of g/cm^3.

1.59. **Collect and Organize**

In this problem we use the mass of a carat (the unit of weight for diamonds) to find the mass of a large diamond and then use the density to calculate the volume of that large diamond.

Analyze

We need the fact that 1 carat = 0.200 g and that the density is defined as mass per volume. To find the volume of the diamond, we can rearrange the density equation to read

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

Solve

The mass of the 5.0-carat diamond is

$$5.0\ \text{carat} \times \frac{0.200\ \text{g}}{1\ \text{carat}} = 1.0\ \text{g}$$

The volume of the diamond is then

$$1.0\ \text{g} \times \frac{1\ \text{cm}^3}{3.51\ \text{g}} = 0.28\ \text{cm}^3$$

Think About It

For this relatively large diamond in terms of carats, the mass is fairly small (one gram is about one-fifth of the mass of a nickel); in this case, even though the density is relatively low, the volume is also quite small.

1.60. **Collect and Organize**

There is a small amount of mercury in this lake per liter, but the volume of the lake will be quite large. In this problem we have to find the volume of the lake and use the concentration of mercury in one liter of the lake water to find the total amount of mercury in the lake.

Analyze

The volume of the lake can be calculated from the surface area and the average depth. However, we need this answer in liters since the concentration of mercury is expressed in micrograms per liter. We should first convert square kilometers to square meters. Then, we can calculate the volume of the lake by multiplying the surface area by the average depth. We need these conversions and the formula for volume from surface area and depth:

$$\frac{1\ \text{km}}{1000\ \text{m}}$$

$$\text{Volume of lake (m}^3\text{)} = \text{surface area (m}^2\text{)} \times \text{depth (m)}$$

We then need to convert cubic meters into liters using

$$\frac{1\ \text{m}^3}{1000\ \text{L}}$$

Next, we can use the mercury concentration to find the total mass of mercury in the lake (which will be in micrograms) and convert that to kilograms:

$$\text{Mass of mercury in lake } (\mu\text{g}) = \text{volume of lake (L)} \times \frac{\text{mass of mercury } (\mu\text{g})}{\text{volume (L)}}$$

$$\text{Mass of mercury (kg)} = \text{mass of mercury } (\mu\text{g}) \times \frac{1\ \text{g}}{1 \times 10^6\ \mu\text{g}} \times \frac{1\ \text{kg}}{1000\ \text{g}}$$

Solve

$$\text{Surface area of lake } (\text{m}^2) = 10.0\ \text{km}^2 \times \left(\frac{1000\ \text{m}}{1\ \text{km}}\right)^2 = 1.00 \times 10^7\ \text{m}^2$$

$$\text{Volume of lake } (\text{m}^3) = \left(1.00 \times 10^7\ \text{m}^2\right) \times 15.0\ \text{m} = 1.50 \times 10^8\ \text{m}^3$$

$$\text{Volume of lake (L)} = 1.50 \times 10^8\ \text{m}^3 \times \frac{1000\ \text{L}}{1\ \text{m}^3} = 1.50 \times 10^{11}\ \text{L}$$

The amount of mercury in the lake then is computed as

$$1.50 \times 10^{11}\ \text{L} \times \frac{0.33\ \mu\text{g}}{1\ \text{L}} \times \frac{1\ \text{g}}{1 \times 10^6\ \mu\text{g}} \times \frac{1\ \text{kg}}{1000\ \text{g}} = 50.\ \text{kg}$$

Think About It

Although the concentration of the mercury is quite low, the entire lake contains a relatively large amount of mercury.

1.61. **Collect and Organize**

In this problem we are asked address the number of data points we can eliminate from a data set by applying Grubbs' test.

Analyze

Grubbs' test allows us to identify outliers, or data points that are far away from the mean of a data set. In order to determine if a data point should be rejected, we must calculate the range of the data and the gap between the questionable point and the nearest value.

Solve

Only one data point may be identified using Grubbs' test.

Think About It

If multiple data points are far away from the mean of a data set, it is implied that a significant error in measurement or experiment has occurred. This data set would not be very precise.

1.62. **Collect and Organize**

We are asked to describe which confidence interval (50%, 90%, or 95%) represents the largest range of data.

Analyze

Confidence intervals represent the range of values around a calculated mean that we expect to contain the true mean value. The confidence interval will be largest when the confidence level is higher, as this will encompass a wider range of data points.

Solve

The 50% confidence interval represents the largest range of data.

Think About It

The confidence interval only describes our certainty that a data point will be near the mean of that data set. The actual value could be within that range or entirely different, depending on the accuracy of our measurements.

1.63. **Collect and Organize**

In this problem we are to determine if the 95% confidence interval or one standard deviation is larger for a data set with seven data points.

Analyze

The equation for determining the 95% confidence interval is

$$\mu = \bar{x} \pm \frac{ts}{\sqrt{n}}$$

For a data set with seven data points ($n = 7$), $t = 2.447$ at the 95% CI. The standard deviation for this sample is given by s in the above equation.

Solve

The 95% confidence interval may be expressed as

$$\mu = \bar{x} \pm \frac{ts}{\sqrt{n}} = \bar{x} \pm \frac{2.447\,s}{\sqrt{7}} = \bar{x} \pm 0.925\,s$$

The range described by the 95% confidence interval is thus smaller than the standard deviation, s.

Think About It

The values of t and n change with the sample size, so for smaller data sets ($n \leq 6$), the standard deviation is larger than the 95% CI.

1.64. **Collect and Organize**

We are asked to consider the conditions for rejecting a statistical outlier at the 90%, 95%, and 99% confidence intervals.

Analyze

A data point must have a *Q* value larger than that listed in the Table 1.8. By looking at Table 1.8, we can see that the values of *Q* for rejecting an outlier are larger for a given sample size as we move from the 90% to 95% to 99% confidence intervals.

Solve

(a) Maybe. If a value cannot be labeled as an outlier at the 95% CI, there is a chance that the *Q* value could still be larger than that required to label it as an outlier at the 90% CI. This depends on the size of the gap between the questionable data point and the nearest data point in the set.

(b) No. If a value cannot be labeled as an outlier at the 95% CI, it will never be larger than the *Q* value at the 99% confidence interval.

Think About It

The experimental value of *Q* for the questionable data point will remain constant at the 90%, 95%, and 99% confidence intervals. The minimum value of *Q* to reject a data point changes with each confidence interval, however.

1.65. **Collect and Organize**

This problem asks which value from the given set contains the fewest number of significant figures.

Analyze

The listed values are a mix of standard and scientific notation. For values expressed using scientific notation, only the values prior to the exponent are counted as significant. By expressing all the quantities using scientific notation, we can better compare the number of significant figures.

(a) 5.45×10^2
(b) 6.4×10^{-3}
(c) 6.50×10^0
(d) 1.346×10^2

Solve

(b) 6.4×10^{-3} contains two significant figures, which is the smallest number of significant figures for this set.

Think About It

The value 6.50 has three significant figures because the zero trailing the decimal is significant.

1.66. **Collect and Organize**

This problem asks which value from the given set contains the greatest number of significant figures.

Analyze

The listed values are a mix of standard and scientific notation. For values expressed using scientific notation, only the values prior to the exponent are counted as significant. By expressing all the quantities using scientific notation, we can better compare the number of significant figures.

(a) 5.45×10^2
(b) 6.4×10^{-3}
(c) 6.50×10^0
(d) 1.346×10^2

Solve
The greatest number of significant figures present is four. The quantity with four significant figures is (d) 1.346×10^2.

Think About It
Significant figures express our confidence in a number.

1.67. **Collect and Organize**
This problem asks which value from the given set contains the fewest number of significant figures.

Analyze
These values are all fractions. We should evaluate each fraction and express it using scientific notation to help determine the number of significant figures.

(a) $1/545 = 0.00183 = 1.83 \times 10^{-3}$
(b) $1/6.4 \times 10^{-3} = 156.25 = 1.6 \times 10^2$
(c) $1/6.50 = 0.1538 = 1.54 \times 10^{-1}$
(d) $1/1.346 \times 10^2 = 0.007429 = 7.429 \times 10^{-3}$

Solve
The smallest number of significant figures present is two. The quantity with two significant figures is (b) $1/6.4 \times 10^{-3}$.

Think About It
The number of significant figures is derived from the denominator of the fraction. Despite the fact that $1/6.4 \times 10^{-3}$ gives us a number that appears to contain at least three digits before the decimal, we cannot be more confident than we would be in 6.4×10^{-3}, which contains two significant figures.

1.68. **Collect and Organize**
This problem asks which value from the given set contains the greatest number of significant figures.

Analyze
These values are all fractions. We should evaluate each fraction and express it using scientific notation to help determine the number of significant figures.

(a) $1/545 = 0.00183 = 1.83 \times 10^{-3}$
(b) $1/6.4 \times 10^{-3} = 156.25 = 1.6 \times 10^2$
(c) $1/6.50 = 0.1538 = 1.54 \times 10^{-1}$
(d) $1/1.346 \times 10^2 = 0.007429 = 7.429 \times 10^{-3}$

Solve
The greatest number of significant figures present is four. The quantity with four significant figures is (d) $1/1.346 \times 10^2$.

Think About It
Despite the large number of digits after the decimal that a calculator can generate, the number of significant figures is determined using the denominator of these fractions.

1.69. **Collect and Organize**

From the values given, we must identify those that contain four significant figures.

Analyze

Writing all the quantities in scientific notation will help determine the number of significant figures in each.

(a) $0.0592 = 5.92 \times 10^{-2}$

(b) $0.08206 = 8.206 \times 10^{-2}$

(c) 8.314

(d) $5420 = 5.42 \times 10^{3}$ or 5.420×10^{3} (if the 0 is significant)

(e) 5.4×10^{3}

Solve

The quantities that have four significant figures are (b) 0.08206, (c) 8.314, and maybe (d) 5420, if the 0 is significant.

Think About It

Remember that zeros at the end of the number may be significant or they may simply be acting as placeholders.

1.70. **Collect and Organize**

From the values given, we must identify those that contain three significant figures.

Analyze

Writing all the quantities in scientific notation will help determine the number of significant figures in each.

(a) 7.02

(b) 6.452

(c) 6.02×10^{23}

(d) $302 = 3.02 \times 10^{2}$

(e) $12.77 = 1.277 \times 10^{1}$

Solve

The quantities that have three significant figures are (a) 7.02, (c) 6.02×10^{23}, and (d) 302.

Think About It

Remember that a zero between two other digits is always significant.

1.71. **Collect and Organize**

We are to express the result of each calculation to the correct number of significant figures.

Analyze

The rules regarding the significant figures that carry over in calculations are given in Section 1.8 in the textbook. Remember to operate on the weak-link principle.

Solve

(a) The least well-known value has three significant figures, so the calculator result of 17.363 is reported as 17.4 with rounding up to the tenths place.

(b) The least well-known value has only one significant figure, so the calculator result of 1.044×10^{-13} is reported as 1×10^{-13}.

(c) The least well-known value has three significant figures, so the calculator result of 5.701×10^{-23} is reported as 5.70×10^{-23}.
(d) The least well-known value has three significant figures, so the calculator result of 3.5837×10^{-3} is reported as 3.58×10^{-3}.

Think About It
Indicating the correct number of significant figures for a calculated value indicates the level of confidence we have in our calculated value. Reporting too many significant figures would indicate a higher level of precision in our number than we actually have.

1.72. **Collect and Organize**
We are to express the result of each calculation to the correct number of significant figures.

Analyze
The rules regarding the significant figures that carry over in calculations are given in Section 1.8 in the textbook. Remember to operate on the weak-link principle.

Solve
(a) The least well-known value has three significant figures, so the calculator result of 1.5506×10^{-1} is reported as 1.55×10^{-1}.
(b) The least well-known value has three significant figures, so the calculator result of 146.3988 is reported as 146.
(c) The least well-known value has four significant figures, so the calculator result of 2.25857×10^{-2} is reported as 2.259×10^{-2}.
(d) The least well-known value has three significant figures, so the calculator result of 3.5700×10^{3} is reported as 3.57×10^{3}.

Think About It
Indicating the correct number of significant figures for a calculated value indicates the level of confidence we have in our calculated value. Reporting too many significant figures would indicate a higher level of precision in our number than we actually have.

1.73. **Collect and Organize**
Given the data from three different manufacturers of circuit boards for copper line widths, we can determine which manufacturers were precise and which were accurate. Using the data provided, we can determine the mean, standard deviation, and 95% confidence intervals for each manufacturer.

Analyze
The sample mean is given by the equation

$$\bar{x} = \frac{\sum_i (x_i)}{n}$$

The standard deviation for each manufacturer may be obtained by setting up a data table like the one in Table 1.5. Standard deviations for each manufacturer can be calculated using the equation

$$s = \sqrt{\frac{\sum_i (x_i - \bar{x})}{n - 1}}$$

The 95% confidence interval is expressed by

$$\mu = \bar{x} \pm \frac{ts}{\sqrt{n}}$$

For each manufacturer, five samples were provided, so $n = 5$, and $t = 2.776$, as shown in Table 1.6.

By first calculating the range of values for the width of the copper lines from each manufacturer, we can see which manufacturer has the highest precision. By comparing measured line widths with the specified width of 0.500 μm, we can assess each manufacturer's accuracy.

Solve

(a) The mean for each of the manufacturers is given by

$$\text{Manufacturer 1} \quad \bar{x} = \frac{(0.512 + 0.508 + 0.516 + 0.504 + 0.513)}{5} = 0.511\ \mu\text{m}$$

$$\text{Manufacturer 2} \quad \bar{x} = \frac{(0.514 + 0.513 + 0.514 + 0.514 + 0.512)}{5} = 0.513\ \mu\text{m}$$

$$\text{Manufacturer 3} \quad \bar{x} = \frac{(0.500 + 0.501 + 0.502 + 0.502 + 0.501)}{5} = 0.501\ \mu\text{m}$$

The standard deviations are
Manufacturer 1

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s
0.512	0.001	$1. \times 10^{-6}$		
0.508	−0.003	$9. \times 10^{-6}$		
0.516	0.005	2.5×10^{-5}		
0.504	−0.007	4.9×10^{-5}		
0.513	0.002	$4. \times 10^{-6}$		
			8.8×10^{-5}	0.005

Manufacturer 2

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s
0.514	0.001	$1. \times 10^{-6}$		
0.513	0.000	0		
0.514	0.001	$1. \times 10^{-6}$		
0.514	0.001	$1. \times 10^{-6}$		
0.512	−0.001	$1. \times 10^{-6}$		
			$4. \times 10^{-6}$	0.001

Manufacturer 3

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s
0.500	–0.001	$1. \times 10^{-6}$		
0.501	0.000	0		
0.502	0.001	$1. \times 10^{-6}$		
0.502	0.001	$1. \times 10^{-6}$		
0.501	0.000	0		
			$3. \times 10^{-6}$	0.001

(b) The 95% confidence interval for each manufacturer is given by

$$\text{Manufacturer 1} \quad \mu = 0.511 \pm \frac{2.776(0.005)}{\sqrt{5}} = 0.511 \pm 0.006$$

$$\text{Manufacturer 2} \quad \mu = 0.513 \pm \frac{2.776(0.001)}{\sqrt{5}} = 0.513 \pm 0.001$$

$$\text{Manufacturer 3} \quad \mu = 0.501 \pm \frac{2.776(0.001)}{\sqrt{5}} = 0.501 \pm 0.001$$

The 95% confidence interval includes 0.500 μm only for Manufacturer 3.

(c) Manufacturer 1 has a range of 0.516 – 0.504 = 0.012 μm; Manufacturer 2 has a range of 0.514 – 0.512 = 0.002 μm; and Manufacturer 3 has a range of 0.502 – 0.500 = 0.002 μm. Manufacturer 3 is both precise and accurate. Manufacturer 2 is precise, but not accurate. Manufacturer 1 is neither precise nor accurate.

Think About It

In electronic circuit boards, the specifications must be very strictly adhered to, in terms of both precision and accuracy. The claim of precision, without claiming accuracy, may mislead buyers.

1.74. **Collect and Organize**

Given the data from five different blood tests, we can determine the mean and standard deviation for the group of individuals whose blood was tested. By calculating the 95% confidence interval for this data set, we can determine if the group as a whole is close to the threshold for being considered diabetic.

Analyze

We can calculate the mean and standard deviation for these samples using the equations below and by setting up a data table like that shown in Table 1.5.

$$\bar{x} = \frac{\sum_i (x_i)}{n}, \quad s = \sqrt{\frac{\sum_i (x_i - \bar{x})^2}{n - 1}}$$

The 95% confidence interval is given by the equation

$$\mu = \bar{x} \pm \frac{ts}{\sqrt{n}}$$

For a sample with five data points ($n = 5$), the value for t is 2.776, as shown in Table 1.6.

Solve

(a) $$\bar{x} = \frac{(106 + 99 + 109 + 108 + 105)}{5} = 105 \text{ mg/dL}$$

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s
106	1	1		
99	−6	36		
109	4	16		
108	3	9		
105	0	0		
			62	3.9

(b) $\mu = 105 \pm \frac{2.776(3.9)}{\sqrt{5}} = 105 \pm 4.8$ mg/dL

120 mg/dL is not contained in the 95% confidence interval for this data set. The maximum value contained within this confidence interval is 109.8 mg/dL.

Think About It

Notice that we cannot correctly calculate the 95% confidence interval with incorrect values for the mean and standard deviation.

1.75. **Collect and Organize**

This problem asks us to consider a data set consisting of six measurements and asks us to determine if the largest value is far enough away from the rest of the data points that it should be considered a statistical outlier and discarded.

Analyze

To determine if an outlier is present, we must first calculate the mean and standard deviation of the data set, then calculate the *Z* value. Comparing this *Z* value to the listed reference *Z* values will allow us to determine if the data point is an outlier. The sample mean is given by the equation

$$\bar{x} = \frac{\sum_i (x_i)}{n}$$

The standard deviation may be obtained by setting up a data table like the one in Table 1.5. Standard deviations can be calculated using the equation

$$s = \sqrt{\frac{\sum_i (x_i - \bar{x})^2}{n - 1}}$$

The equation for calculating the *Z* value (Grubbs' test) is

$$Z = \frac{|x_i - \bar{x}|}{s}$$

Six data points were provided, so $n = 6$, and the reference *Z* value from Table 1.8 is 1.887 at the 95% confidence level. If our calculated *Z* value is larger than the reference *Z* value, the data point is in fact an outlier.

Solve

The mean for this data set is

$$\bar{x} = \frac{(3.15 + 3.03 + 3.09 + 3.11 + 3.12 + 3.41)}{6} = 3.15$$

The standard deviation is

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s

3.15	1	0		
3.03	–0.12	0.014		
3.09	–0.06	0.004		
3.11	–0.04	0.002		
3.12	–0.03	0.0009		
3.41	0.26	0.068		
			0.089	0.13

The largest outlier for this dataset is 3.41. For this data point

$$Z = \frac{0.26}{0.13} = 2.0$$

This Z value is greater than the reference Z value of 1.887, so this data point is in fact an outlier and should be discarded.

Think About It

Even though a data point may appear quite different from the rest of the set, a Grubbs' test should be conducted to determine if it may be rejected. In this case, the Z value was relatively close to the minimum allowable value to discard a point, so a slightly smaller measurement might be valid!

1.76. **Collect and Organize**

This problem asks us to consider a data set consisting of six measurements and asks us to determine if the largest value is far enough away from the rest of the data points that it should be considered a statistical outlier and discarded.

Analyze

To determine if an outlier is present, we must first calculate the mean and standard deviation of the data set, then calculate the Z value. Comparing this Z value to the listed reference Z values will allow us to determine if the data point is an outlier. The sample mean is given by the equation

$$\bar{x} = \frac{\sum_i (x_i)}{n}$$

The standard deviation may be obtained by setting up a data table like the one in Table 1.5. Standard deviations can be calculated using the equation

$$s = \sqrt{\frac{\sum_i (x_i - \bar{x})^2}{n-1}}$$

The equation for calculating the Z value (Grubbs' test) is

$$Z = \frac{|x_i - \bar{x}|}{s}$$

Six data points were provided, so $n = 6$, and the reference Z value from Table 1.8 is 1.887 at the 95% confidence level. If our calculated Z value is larger than the reference Z value, the data point is in fact an outlier.

Solve

The mean for this data set is

$$\bar{x} = \frac{(61 + 75 + 64 + 65 + 64 + 66)}{6} = 66$$

The standard deviation is

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s

61	–5	25		
75	9	81		
64	–2	4		
65	–1	1		
64	–2	4		
66	0	0		
			115	4.80

The largest outlier for this dataset is 75. For this data point

$$Z = \frac{9}{4.80} = 1.88$$

This Z value is approximately equal to the reference Z value of 1.887, so this data point might be an outlier.

Think About It

Since the Z value is approximately the same as the reference Z value for $n = 6$ at the 95% confidence level, it is about 95% likely that the data point is a statistical outlier.

1.77. **Collect and Organize**

This question asks if a temperature in Celsius would ever equal the temperature in Fahrenheit. We have to make use of the conversion equation between Celsius and Fahrenheit degrees.

Analyze

The equation converting between the temperatures is given as

$$^\circ\text{C} = \frac{5}{9}(^\circ\text{F} - 32)$$

To find the temperature at which these temperature scales meet (°C = °F in the above equation) substitute °C for °F:

$$^\circ\text{C} = \frac{5}{9}(^\circ\text{C} - 32)$$

Solve

Rearranging this equation and solving for °C gives

$$^\circ\text{C} = \frac{5}{9}(^\circ\text{C} - 32)$$

$$\frac{9}{5}(^\circ\text{C}) = (^\circ\text{C} - 32)$$

$$\frac{9}{5}(^\circ\text{C}) - {}^\circ\text{C} = -32$$

$$\frac{4}{5}(^\circ\text{C}) = -32$$

$$^\circ\text{C} = -40^\circ\text{C} = -40^\circ\text{F}$$

Think About It
Because the intervals between degrees on the Celsius scale are larger than the degrees on the Fahrenheit scale, the two scales will eventually meet at one temperature. This solution shows that they meet at –40°.

1.78. **Collect and Organize**
In this question we define the *absolute* temperature scale.

Analyze
The Kelvin scale is the absolute temperature scale, and its lowest temperature is 0 K.

Solve
The absolute temperature scale (Kelvin scale) has no negative temperatures, and its 0 value is placed at the lowest possible temperature.

Think About It
Because the Kelvin scale has no negative temperatures, it will often be used in equations when using a negative temperature (in Celsius) would result in a nonsensical answer.

1.79. **Collect and Organize**
Given the freezing and boiling point of a radiator coolant, we are to convert these temperatures from the Celsius scale to the Fahrenheit scale.

Analyze
The relationship between Celsius and Fahrenheit temperature scales is given by

$$°\text{C} = \frac{5}{9}(°\text{F} - 32)$$

This will have to be rearranged to find °F from °C.

$$°\text{F} = \frac{5}{9}(°\text{C}) + 32$$

Solve
The freezing point of this coolant in degrees Fahrenheit is

$$°\text{F} = \frac{9}{5}(-39.0°\text{C}) + 32 = -38°\text{F}$$

The boiling point of this coolant in degrees Fahrenheit is

$$°\text{F} = \frac{9}{5}(110°\text{C}) + 32 = 230°\text{F}$$

Think About It
In computing the freezing point of this coolant in degrees Fahrenheit, notice that the result is nearly the same as the freezing point in degrees Celsius. This is because the two temperature scales do share a temperature value (see P1.80).

1.80. **Collect and Organize**

This question asks for the conversion of two melting points from Celsius degrees to Kelvin.

Analyze

We need only adjust the values for the temperature according to the following equation:

$$K = °C + 273.15$$

Solve

For the melting point of silver

$$K = 962°C + 273.15 = 1235 \text{ K}$$

For the melting point of gold

$$K = 1064°C + 273.15 = 1337 \text{ K}$$

Think About It

The two melting temperatures should remain the same number of units apart (102) because the Kelvin is the same magnitude as the Celsius degree.

1.81. **Collect and Organize**

We are asked to compare the critical temperature (T_c) of three superconductors. The critical temperatures, however, are given in three different temperature scales, so for the comparison, we will need to convert them to a single scale.

Analyze

It does not matter which temperature scale we use as the common one, but since the critical temperatures are low, it might be easiest to express all of the temperatures in Kelvin. The equations we will need are

$$K = °C + 273.15 \quad \text{and} \quad °C = \frac{5}{9}(°F - 32)$$

Solve

The T_c for the first material is already expressed in Kelvin, $T_c = 93.0$ K.
The T_c of the second material is expressed in °C and can be converted to K by

$$K = -250.0°C + 273.15 = 23.2 \text{ K}$$

The T_c of the third material is expressed in Fahrenheit degrees. To get this temperature in Kelvin, first convert to Celsius degrees, then to Kelvin:

$$°C = \frac{5}{9}(-231°F - 32) = -146.2°C$$

$$K = -146.2°C + 273.15 = 127.0 \text{ K}$$

The superconductor with the highest T_c is expressed in °F with a T_c of 127.0 K.

Think About It

The superconductor with the lowest T_c is the second material with a T_c of 23.2 K, more than 100 K lower than the T_c of that with the highest T_c.

1.82. **Collect and Organize**

Based on the boiling points for N_2, O_2, and Ar, we can determine which gas condenses first as air is cooled.

Analyze
The boiling point of the liquids from Appendix 3, Table A3.2 are as follows:

$$\text{b.p. } N_2 = -195.8°C$$
$$\text{b.p. } O_2 = -182.95°C$$
$$\text{b.p. Ar} = -185.9°C$$

Solve
As air is cooled, the gas with the highest (least negative) boiling point will condense first. Oxygen, therefore, will condense first.

Think About It
The second gas to condense is argon, followed by nitrogen.

1.83. **Collect and Organize**
We are asked in this problem to convert from Kelvin to degrees Celsius.

Analyze
The relationship between the Kelvin temperature scale and the Celsius temperature scale is given by

$$K = °C + 273.15$$

Rearranging this gives the equation to convert Kelvin temperatures to Celsius:

$$°C = K - 273.15$$

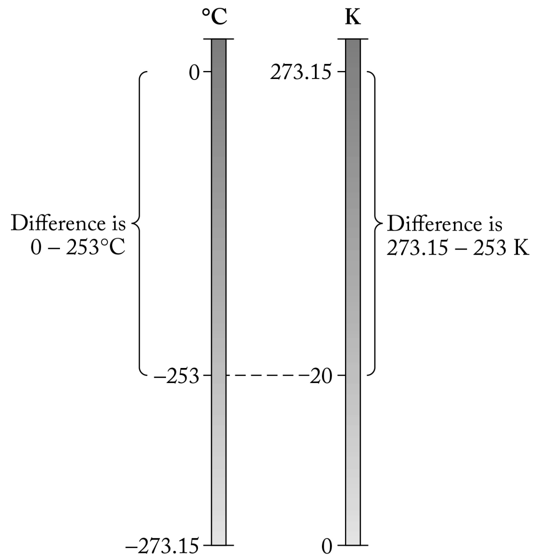
Solve

$$°C = 4.2\ K - 273.15 = -269.0°C$$

Think About It
Because 4.2 K is very cold, we would expect that the Celsius temperature would be very negative. It should not, however, be lower than −273.15°C since that is the lowest temperature possible.

1.84. **Collect and Organize**
This question asks us to convert a Celsius temperature into Kelvin.

Analyze
Though temperatures on each scale have different values, the difference between temperatures on the Kelvin and Celsius scales is the same. We can easily convert between scales by adding or subtracting the difference between the two scales at absolute zero, as shown in the figure:



The relationship between the Kelvin temperature scale and the Celsius temperature scale is given by

$$K = °C + 273.15$$

Solve

$$K = -253°C + 273.15 = 20.\ K$$

Think About It

This is a low temperature for the Celsius temperature scale. It is still above the lowest possible temperature (0 K or –273.15°C); it has to be!

1.85. **Collect and Organize**

This question asks us to convert a temperature in Fahrenheit degrees to a temperature in Celsius degrees.

Analyze

The relationship between the Celsius and Fahrenheit temperature scales is given by

$$°C = \frac{5}{9}(°F - 32)$$

Solve

$$°C = \frac{5}{9}(102.5°F - 32) = 39.2°C$$

Think About It

This temperature makes sense because it is higher than the normal body temperature, 37°C.

1.86. **Collect and Organize**

This question asks us to convert a temperature in Celsius degrees to a temperature in Fahrenheit degrees.

Analyze

The relationship between the Celsius and Fahrenheit temperature scales is

$$°C = \frac{5}{9}(°F - 32)$$

This will have to be rearranged to find °F from °C:

$$°F = \frac{9}{5}(°C) + 32$$

Solve

$$°F = \frac{9}{5}(37.0°C) + 32 = 98.6°F$$

Think About It

In humans, this temperature is considered normal body temperature.

1.87. **Collect and Organize**

This question asks us to convert the coldest temperature recorded on Earth from Fahrenheit to Celsius and Kelvin.

Analyze

Since the Celsius and Kelvin scales are similar (offset by 273.15 degrees), once we convert from Fahrenheit to Celsius, finding the Kelvin temperature will be straightforward. The equations we need are

$$°C = \frac{5}{9}(°F - 32)$$

$$K = °C + 273.15$$

Solve

$$°C = \frac{5}{9}(-128.6°F - 32) = -89°C$$

$$K = -89.2°C + 273.15 = 183.9\,K$$

Think About It

This temperature is cold on any scale!

1.88. **Collect and Organize**

This question asks us to convert the hottest temperature recorded on Earth from Fahrenheit to Celsius and Kelvin.

Analyze

Since the Celsius and Kelvin scales are similar (offset by 273.15 degrees), once we convert from Fahrenheit to Celsius, finding the Kelvin temperature will be straightforward. The equations we need are

$$°C = \frac{5}{9}(°F - 32)$$

$$K = °C + 273.15$$

Solve

$$°C = \frac{5}{9}(134°F - 32) = 56.7°C$$

$$K = 56.7°C + 273.15 = 329.8\ K$$

Think About It

Remind yourself about the expected range for temperatures in the Celsius scale. While 56.7°C is quite warm for a human, it is still below the boiling point of water; a cup of coffee at this temperature would seem tepid and unpalatable!

1.89. **Collect and Organize**

In this problem we need to express each mixture of chlorine and sodium as a ratio. The mixture that is closest to the ratio for chlorine to sodium will be the one with the desired product, leaving neither sodium nor chlorine left over.

Analyze

First, we must calculate the ratio of chlorine to sodium in sodium chloride. This is a simple ratio of the masses of these two substances:

$$\frac{\text{mass of chlorine}}{\text{mass of sodium}} = \text{ratio of the two components}$$

We can compare the ratios of the other mixtures by making the same calculations.

Solve

In sodium chloride the mass ratio of chlorine to sodium is

$$\frac{1.54\ \text{g chlorine}}{1.00\ \text{g sodium}} = 1.54$$

Repeating this calculation for the four mixtures, we obtain each ratio of chlorine to sodium:

$$\frac{17.0\ \text{g}}{11.0\ \text{g}} = 1.55 \text{ for mixture a} \qquad \frac{12.0\ \text{g}}{6.5\ \text{g}} = 1.8 \text{ for mixture c}$$

$$\frac{10.0\ \text{g}}{6.5\ \text{g}} = 1.5 \text{ for mixture b} \qquad \frac{8.0\ \text{g}}{6.5\ \text{g}} = 1.2 \text{ for mixture d}$$

Both mixtures a and b react so that there is neither sodium nor chlorine left over.

Think About It

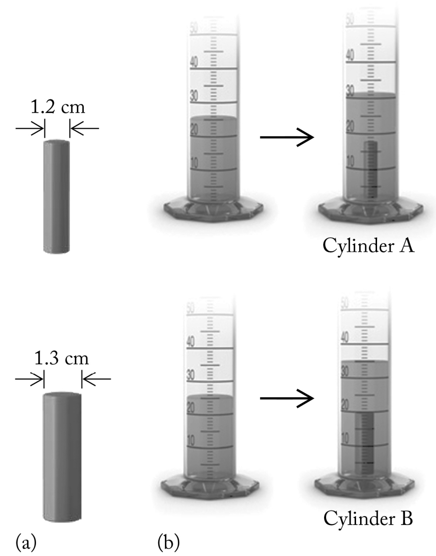
Mixture c has leftover chlorine and mixture d has leftover sodium after the reaction is complete.

1.90. **Collect and Organize**

For this problem we try to identify which cylinder is made of aluminum and which is made of titanium by comparing experimentally determined densities with the actual known densities.

Analyze

It may help to picture what we are asked to determine in each part.



(a) To calculate the volume of each cylinder from its dimensions, we will have to use the equation for the volume of a cylinder:

$$\text{Volume of cylinder} = \text{height of cylinder} \times \pi \times (\text{radius})^2$$

where radius $= \frac{1}{2} \times$ diameter.

(b) To calculate the volume from the water displacement method, we need only find the difference in water volume for each cylinder from the diagram in Figure P1.90.

(c) To determine the method with the most significant figures, we will compare the answers in parts (a) and (b).

(d) To compute the density for each cylinder we use the equation for density:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

Solve

(a) Volumes of the cylinders from their measured dimensions:

$$\text{Volume of Cylinder A} = 5.1\text{ cm} \times \pi \times (0.60\text{ cm})^2 = 5.8\text{ cm}^3$$
$$\text{Volume of Cylinder B} = 5.9\text{ cm} \times \pi \times (0.65\text{ cm})^2 = 7.8\text{ cm}^3$$

(b) Volume of cylinders from water displacement measurements:

$$\text{Volume of Cylinder A} = 30.8\text{ mL} - 25.0\text{ mL} = 5.8\text{ mL}$$
$$\text{Volume of Cylinder B} = 32.8\text{ mL} - 25.0\text{ mL} = 7.8\text{ mL}$$

(c) Neither. As seen in the above calculations, both the volume measurement from the water displacement method and from the cylinders' dimensions have two significant figures for the volume calculation.

(d) From part (a) and part (b) we obtain answers with two significant figures for the density:

$$\text{Density of Cylinder A} = \frac{15.560\text{ g}}{5.8\text{ mL}} = 2.7\text{ g/mL}$$

$$\text{Density of Cylinder B} = \frac{35.536\text{ g}}{7.8\text{ mL}} = 4.6\text{ g/mL}$$

Think About It

How we make measurements is very important to the values we can report for those measurements. In this problem, neither method provided more significant figures for the calculation. By comparing the calculated densities to known values, we see that cylinder A is aluminum, and cylinder B is titanium.

1.91. **Collect and Organize**

This problem asks us to compute the percentages of the two ingredients in trail mix as manufactured on different days. Given these percentages, we are asked to determine the mean, standard deviation, and 90% confidence intervals for each component. With this data, we can determine if the trail mix contains the ideal composition of 67% peanuts and 33% raisins.

Analyze

Each day has a total of 82 pieces (peanuts plus raisins), so the percentage of the mix in peanuts for each day is calculated by the equation

$$\%\text{ peanuts} = \frac{\text{number of peanuts in mix}}{82} \times 100$$

Since the trail mix contains only two components, the percentage of raisins is simply 100% – % peanuts. The mean and standard deviation may be calculated using the following equations:

$$\bar{x} = \frac{\sum_i (x_i)}{n}, \quad s = \sqrt{\frac{\sum_i (x_i - \bar{x})^2}{n-1}}$$

The 90% confidence interval is given by the equation

$$\mu = \bar{x} \pm \frac{ts}{\sqrt{n}}$$

For a sample with four data points ($n = 4$), the value for t is 2.353, as shown in Table 1.6.

Solve

(a) For each day, the percentage of peanuts and raisins is

$$\text{Day 1} \quad \frac{50\text{ peanuts}}{82\text{ pieces}} \times 100\% = 61\%\text{ peanuts}$$

$$100\% - 61\% = 39\%\text{ raisins}$$

$$\text{Day 11} \quad \frac{56\text{ peanuts}}{82\text{ pieces}} \times 100\% = 68\%\text{ peanuts}$$

$$100\% - 68\% = 32\%\text{ raisins}$$

$$\text{Day 21} \quad \frac{48\text{ peanuts}}{82\text{ pieces}} \times 100\% = 59\%\text{ peanuts}$$

$$100\% - 59\% = 41\%\text{ raisins}$$

$$\text{Day 31} \quad \frac{52\text{ peanuts}}{82\text{ pieces}} \times 100\% = 63\%\text{ peanuts}$$

$$100\% - 63\% = 37\%\text{ raisins}$$

We may calculate the mean and standard deviation for each component.

$$\text{Peanuts} \quad \bar{x} = \frac{(61 + 68 + 59 + 63)}{4} = 63\%$$

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s
61	–2	4		
68	5	25		
59	–4	16		
63	0	0		
			45	3.9

$$\text{Raisins} \quad \bar{x} = \frac{(39 + 32 + 41 + 37)}{4} = 37\%$$

x_i	$(x_i - \bar{x})$	$(x_i - \bar{x})^2$	$\sum_i (x_i - \bar{x})^2$	s
39	–2	4		
32	5	25		
41	–4	16		
37	0	0		
			45	3.9

(b)

$$\text{Peanuts} \quad \mu = 63 \pm \frac{2.353(3.9)}{\sqrt{4}} = 63 \pm 4.6\%$$

$$\text{Raisins} \quad \mu = 37 \pm \frac{2.353(3.9)}{\sqrt{4}} = 37 \pm 4.6\%$$

Yes, the 90% confidence interval contains the percentages for the desired composition of 67% peanuts and 33% raisins.

Think About It

On Days 1, 21, and 31, too few peanuts were in the trail mix.

1.92. **Collect and Organize**

For this problem we have to calculate the volume that each liquid takes up from their known density values and add those volumes together. We then have to use that volume to determine the height of the liquid in the cylinder.

Analyze

To find the volume of each liquid from the given mass we can use the density equation:

$$\text{Volume} = \frac{\text{mass}}{\text{density}}$$

The volume of each liquid should then be added together to find the total volume. Rearranging the equation describing the volume of a cylinder as shown, we can calculate the height of the two liquids in the cylinder:

$$\text{Volume of cylinder } (V) = \pi r^2 h$$

$$h = \frac{V}{\pi r^2}$$

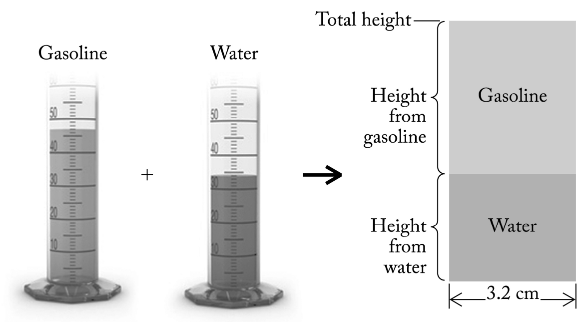
We have to be careful to watch for consistent units. Knowing that 1 cm^3 = 1 mL should help.

Solve

$$\text{Volume of gasoline} = 34.0\text{ g} \times \frac{1\text{ mL}}{0.73\text{ g}} = 47\text{ mL}$$

$$\text{Volume of water} = 34.0\text{ g} \times \frac{1\text{ mL}}{1.00\text{ g}} = 34.0\text{ mL}$$

$$\text{Total volume} = 47\text{ mL} + 34.0\text{ mL} = 81\text{ mL}$$



Using the above equation with V = 81 mL or 81 cm^3 and the fact that the radius is half the diameter (3.2 cm/2 = 1.6 cm) gives the height of the liquids in the cylinder.

$$h = \frac{81\text{ cm}^3}{\pi \times (1.6\text{ cm})^2} = 10.\text{ cm}$$

Think About It

This is a reasonable value for the height of the liquids in the cylinder considering that their combined volume is about 81 mL.

1.93. **Collect and Organize**

We are told that the same force is used to stretch two springs. The stronger spring will stretch less than the weaker spring, so we are looking for which spring will experience a smaller percentage increase in length.

Analyze

We can find the percentage length increase in spring A by considering the length when stretched ($l_{\text{stretched}}$) and the natural length (l_{natural}). We can determine the percentage increase in length using

$$\%\text{ length increase} = \frac{l_{\text{stretched}} - l_{\text{natural}}}{l_{\text{natural}}} \times 100\%$$

Solve

$$\%\text{ length increase} = \frac{5.4\text{ cm} - 4.0\text{ cm}}{4.0\text{ cm}} \times 100\% = 35\%$$

Spring A’s length increases by 35%, while spring B’s length only increases by 15%. Spring B is stronger.

Think About It

If we assume that spring B was also 4.0 cm in its natural state, a 15% increase in length corresponds to a final length of 4.6 cm. This is much shorter than spring A when stretched, supporting our answer above.

1.94. **Collect and Organize**

This question asks us about a total distance traveled in both miles and kilometers, given a time and a series of speeds. A speed expressed as 1.6 mi/h is equivalent to saying "travels 1.6 miles in one hour."

Analyze

We can multiply each speed by the time the class traveled each day, eight hours, to determine the number of miles traveled each day. By adding these numbers, we can determine the total distance traveled on the field trip. The daily speeds we need are given in the question:

$$\frac{1.6\text{ mi}}{1\text{ h}}, \frac{1.4\text{ mi}}{1\text{ h}}, \frac{1.7\text{ mi}}{1\text{ h}}$$

Solve

$$\left(8\text{ h}\times\frac{1.6\text{ mi}}{1\text{ h}}\right)+\left(8\text{ h}\times\frac{1.4\text{ mi}}{1\text{ h}}\right)+\left(8\text{ h}\times\frac{1.7\text{ mi}}{1\text{ h}}\right)=38\text{ mi}$$

$$\left[\left(8\text{ h}\times\frac{1.6\text{ mi}}{1\text{ h}}\right)+\left(8\text{ h}\times\frac{1.4\text{ mi}}{1\text{ h}}\right)+\left(8\text{ h}\times\frac{1.7\text{ mi}}{1\text{ h}}\right)\right]\times\frac{1.6093\text{ km}}{1\text{ mi}}=61\text{ km}$$

Think About It

For a three-day hike, 38 miles seems like a reasonable distance. The class hikes nearly a half-marathon (13.1 miles) each day!

1.95. **Collect and Organize**

From the given mass of fluoride per gram of toothpaste and given that NaF is 45% fluoride by mass, we are to calculate the mass of NaF in an 8.2-oz tube of toothpaste.

Analyze

First, we will convert the weight of the toothpaste in the tube from ounces to grams using 1 oz = 28.35 g. From that we can calculate the mg of F^- in the toothpaste using 1.00 mg F^-/g of toothpaste. Finally, because 45% of the mass of the active ingredient in toothpaste (NaF) is F^-, we multiply the result by 100/45 to obtain the mass of sodium fluoride (NaF) in the toothpaste.

Solve

$$8.2\text{ oz}\times\frac{28.35\text{ g toothpaste}}{1\text{ oz}}\times\frac{1.00\text{ mg F}^-}{1\text{ g toothpaste}}\times\frac{100\text{ mg NaF}}{45\text{ mg F}^-}=5.2\times10^2\text{ mg NaF}$$

Think About It

Some toothpastes use sodium monofluorophosphate (Na_2PO_3F) as the active ingredient instead of sodium fluoride. Sodium monofluorophosphate has a much smaller percentage of fluoride, so a larger mass is required to deliver the same quantity of fluoride in toothpaste.

1.96. **Collect and Organize**

Given the amounts of recognition molecule, capture molecule, and detector molecule along with the amounts of each needed for one HIV assay plate, we are to determine whether we have sufficient amounts for 96 assay plates.

Analyze
To calculate the amount of each molecule required for the 96 assay plates, we first multiply the required amounts of each molecule by 96. We then subtract this amount from the quantity of each molecule that is available to determine if there are sufficient amounts of each molecule.

Solve
(a) Amount of recognition molecule needed:

$$0.550\ \text{mg} \times 96 = 52.8\ \text{mg}$$

This is less than the amount of recognition molecule available.

$$100.00\ \text{mg} - 52.8\ \text{mg} = 47.2\ \text{mg left over}$$

Amount of capture molecule needed:

$$1.200\ \text{mg} \times 96 = 115.2\ \text{mg}$$

This is more than the amount of recognition molecule available.

$$100.00\ \text{mg} - 115.2\ \text{mg} = -15.2\ \text{mg (deficit)}$$

Amount of detector molecule needed:

$$0.450\ \text{mg} \times 96 = 43.2\ \text{mg}$$

This is less than the amount of detector molecule available.

$$50.00\ \text{mg} - 43.2\ \text{mg} = 6.8\ \text{mg left over}$$

(b) We can make only the number of plates that use up the amount of capture molecules.

$$\frac{100.00\ \text{mg}}{1.200\ \text{mg}} = 83\ \text{plates}$$

Think About It
At most we could prepare 83 plates. Practically speaking, however, we might expect that we would be able to prepare slightly fewer than 83 plates because it is likely that some material would be lost during weighing and transfer of the capture molecule sample.

1.97. **Collect and Organize**

In this problem we have to calculate the number of tablets of vitamin C to take to reach a dosage level of 1.00 g given the amount of vitamin C in a 500.0 mg tablet.

Analyze
First, we have to calculate the amount (in grams) of vitamin C in each 500.0 mg tablet:

$$\text{Mass of vitamin C in tablet} = \text{mass of tablet}' \ \text{percentage of vitamin C in tablet}$$

Next, we determine the number of tablets that would comprise 1.00 g of vitamin C:

$$\text{Number of tablets} = \frac{\text{amount of vitamin C desired}}{\text{amount of vitamin C in one tablet}}$$

Solve

$$500.0\ \text{mg}' \ \frac{1\ \text{g}}{1000\ \text{mg}}' \ 0.20 = 0.10\ \text{g of vitamin C in each tablet}$$

$$\frac{1.00 \text{ g vitamin C}}{0.10 \text{ g of vitamin C/tablet}} = 10 \text{ tablets}$$

Think About It

This does not seem to be an extraordinary number of tablets to take. This is because the tablets themselves are fairly large (0.5 g) and contain a good portion of vitamin C (20%).

1.98. **Collect and Organize**

Given the low and high measured temperatures for three thermometers, we can choose the best one to use.

Analyze

The best thermometer will be the one that is most accurate for the range of temperatures.

Solve

Thermometer A deviates from 0°C by –0.8°C and from 100°C by +0.1°C. Thermometer B deviates from 0°C by +0.3°C and from 100°C by –0.2°C. Thermometer C deviates from 0°C by +0.3°C and from 100°C by +0.3°C. Only for Thermometer C is the temperature scale linear, as it deviates from the melting and boiling points of water by the same amount (–0.3°C), so it will be the "best" thermometer to use because it will be the most precise.

Think About It

To accurately use Thermometer C you must subtract 0.3°C from each measured value.