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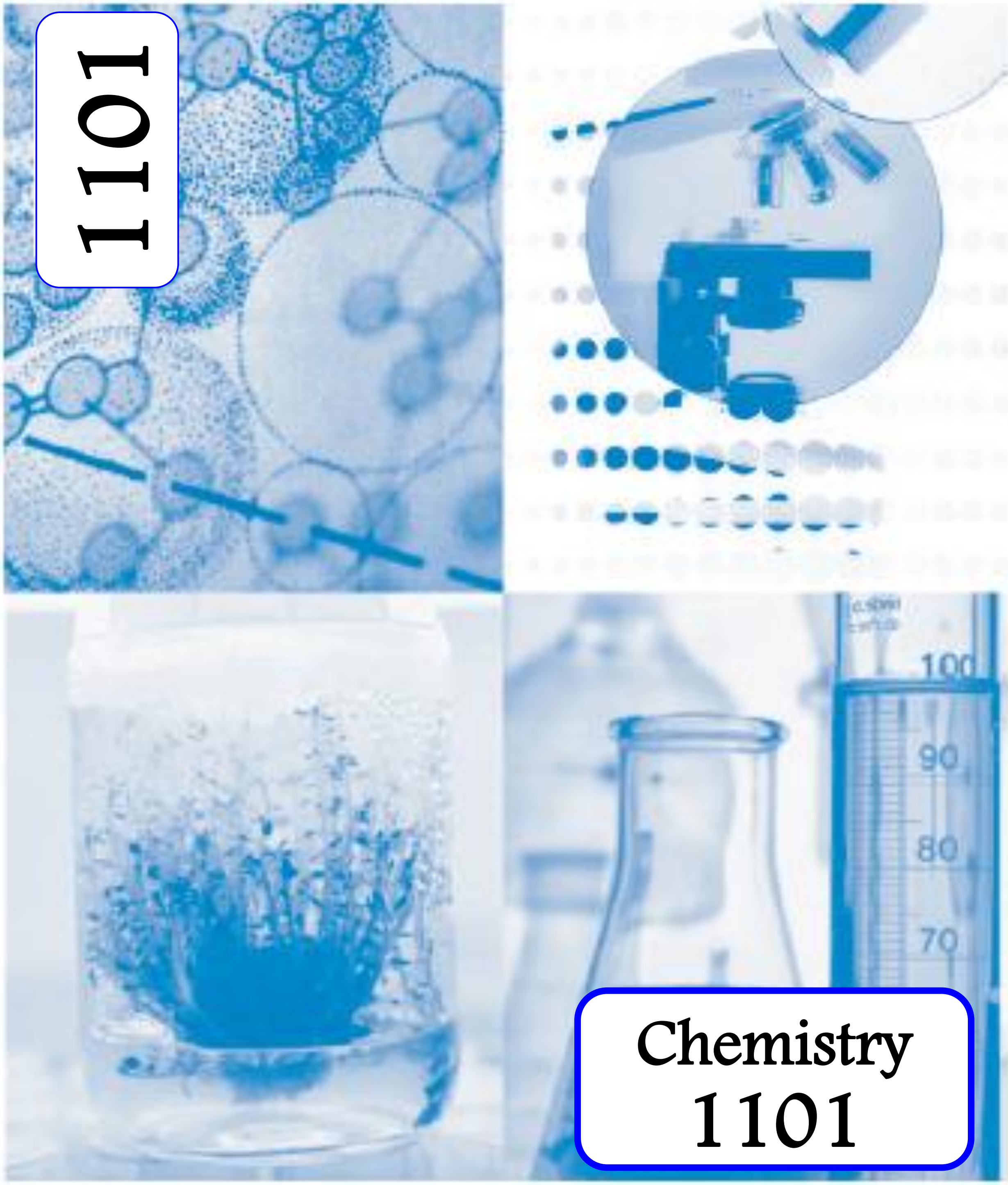

Chemistry 1101

Date: ______________________________________

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CHEMISTRY 1101

Terms, Units and the Scientific Method

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Terms, Units and the Scientific Method

In this unit

Chapter one introduces the scientific method. By the end of this chapter, the student should have a better understanding of the methods used for measuring matter and energy.

The phases of matter are taught as well as an introduction to thermodynamics. The student will be taught the definition of an atom, an element, and a chemical bond.

Glossary of Terms

Accuracy is the agreement between the measurement and the accepted or true value.

Atom: The smallest part of an element that can exist chemically.

Avagadro's constant: Formerly known as Avagadro's number, it s defined as the number of atoms or molecules in one *mole* of the substance.

Chemical bond: An attractive force that holds atoms or ions together.

Chemistry: The branch of science that deals with the materials that make up the universe and the changes that the materials experience. More specifically chemistry deals with the elements and the compounds that they form.

Element: A substance that cannot be decomposed into simpler substances.

Energy: A measure of a system's ability to do work. When something possesses energy, it can affect other objects by performing work on them.

Gas: A gas has no fixed volume or shape. A gas will evenly fill the volume and shape of its container.

Inertia: The property of a body by which it remains at rest or continues moving in a straight line unless acted upon by a directional force.

Kinetic energy: Energy that is a result of the motion of an object and depends on the mass and velocity of the object.

Liquid: A substance that has a definite volume but takes the shape of its container.

Mass: The property of an object that is a measure of its *inertia*. The amount of matter an object contains.

Matter: That which occupies space and has mass.

Mixtures: Two or more substances, which are variable in composition.

Molecule: The smallest part of a compound having all the properties of the compound.

Natural law: A descriptive principle that holds true in all circumstances covered by that law.

Potential energy: Energy that is a result of an object's position or composition.

Precision: Refers to the agreement of two or more measurements of a value.

Solid: A substance that when acted upon by a force tends to maintain both volume and shape.

Standard deviation: Measures how closely the data (the numbers) are clustered about the mean.

Temperature: The measurement of heat.

Theory: An attempt to tell us why something happens.

Thermodynamics: The study of the laws that govern the conversion of energy from one form to another, the direction in which heat will flow, and the availability of energy to do work.

Units: The scale, standard or type of object that is being measured.

Weight: The force upon something that is caused by gravity.

Chapter One

What is Chemistry?

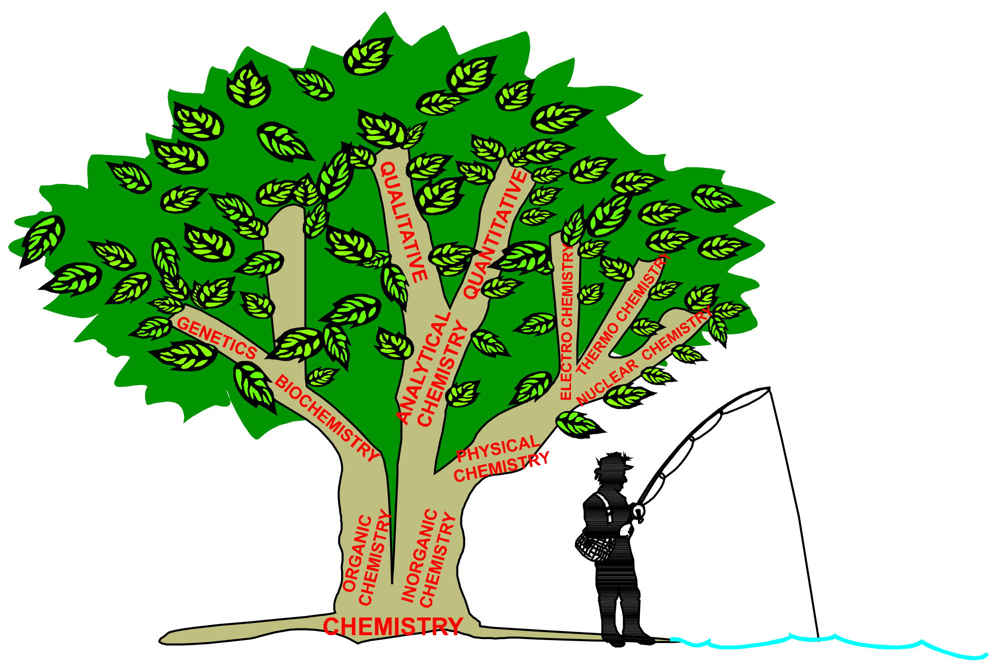
What do you think of when you think about chemistry? Explosions? Exotic drugs? Dangerous toxins? Good stuff, bad or both? Chemistry is neither inherently good nor bad. What we do with chemistry determines its impact upon humankind and the world.

All science is composed of two main components, a *body of knowledge* and a *method of study*. ***Chemistry*** *is the branch of science that deals with the materials that make up the universe and the changes that the materials experience*. More specifically chemistry deals with the elements and the compounds that they form. Chemistry is a systemized method of gaining knowledge through observation and specific experimentation. The study of chemistry enables one to understand better the nature of the matter that surrounds them. The study of chemistry can be fun, try to apply the principles to everyday life, because chemistry is all around us and inside us as well.

Branches of Chemistry

There are two primary categories of science; the *biological* sciences, those sciences that are concerned with living organisms; and the *physical* sciences, those sciences that are concerned with the study of matter in nature. This course primarily covers general chemistry. General Chemistry is a physical science because it is primarily concerned with the study of materials in nature, in the *physical* world.

Chemistry can be divided into branches according to either the substances studied or the types of study conducted. The primary division of the branch that deals with substances is between *inorganic* chemistry and *organic* chemistry. This division exists primarily because the early chemists knew that there was a vast difference between the chemistry of artificially produced compounds (inorganic chemistry) and the chemistry of naturally occurring compounds (organic chemistry).



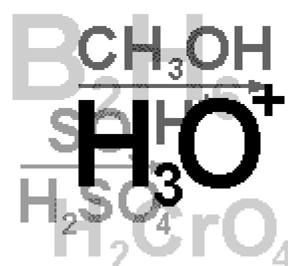
NUCLEAR FISHIN'

UNDER THE Ol' CHEMIS-TREE

Organic chemistry is the study of carbon compounds. At one time scientists thought that carbon compounds could only come from living sources. Then, in the nineteenth century, many organic compounds were artificially produced. Still, the compounds of carbon constitute the central chemicals of all living things on this planet. Organic compounds include small molecules such as methane (CH_4), and giant “macromolecules” such as the deoxyribonucleic acids (DNA), which would be about two meters long if uncoiled.

Organic compounds range from substances such as diamonds, to alcohols, oils, amino acids, celluloses, sugars, plastics, and hemoglobin. Carbon compounds make up the many enzymes that catalyze the reactions that occur in our bodies. There are about 6 million known organic compounds. The special way in which carbon can bond allows for an almost unlimited number of possible compounds. We will study organic chemistry further in Unit Nine (Science 1109).

Inorganic chemistry is the study of the compounds that are not classified as organic. While the study of inorganic chemistry may sometimes include a compound that has carbon in it, it is not the carbon molecule that is being studied. Inorganic chemistry is the study of all the elements and their compounds. The development of the periodic table is one of the great scientific achievements of civilization. The study of inorganic chemistry is the study of the periodic table and relationships of elements and compounds. There are about 200,000 known inorganic compounds. We will study the periodic table in Unit Two (Science 1102).



Physical Chemistry is the branch of inorganic chemistry that is concerned with the physical properties of compounds. This field of science applies the principles of physics to chemical compounds. These physical properties may include electromagnetic properties, surface tension, viscosity (a substance’s resistance to flow), density, vapor pressure, and crystallography (the study of crystal form). Physical chemistry is further divided into electrochemistry, chemical kinetics, and thermochemistry. Thermochemistry is the investigation of the changes in energy and entropy (the measure of disorder) that occur during chemical reactions and in phase changes. Electrochemistry concerns the effect of changes to chemical compounds and chemical substances in the presence of electricity. Chemical kinetics analyzes the details of the equilibrium states between chemical reactants and chemical products. We will take a deeper look at thermodynamics and physical chemistry in Unit Eight (Science 1108).

Analytical chemistry is the separation and identification of materials. It is a collection of techniques that allow the chemist to determine the exact chemical composition of substances. Analytical chemistry is further divided into quantitative and qualitative chemistry. Qualitative chemistry is concerned with the identification of each molecule and compound present (its “quality”). In quantitative chemistry the exact amounts (quantities) of each compound and molecule is precisely determined. We will study quantitative and qualitative chemistry in Unit Seven (Science 1107). Stoichiometry (the relative proportions in which elements form compounds or in which substances react) is studied in analytical chemistry. Stoichiometry (and other big words) will be covered in Unit Nine and elsewhere.

Nuclear chemistry is the study of atomic nuclei and the study of the reactions of nuclei with subatomic particles. Most of the compounds studied are radioactive. We will look at nuclear chemistry and quantum theory in Unit Ten (Science 1110).

Biochemistry is the study of the chemistry of living organisms, especially the structure and function of their chemical components. The prefix "bio-" comes from the Latin word meaning "life". Biochemistry is primarily concerned with the reactions and interactions of biomolecules. A thorough understanding of Organic chemistry is helpful in the study of biochemistry. In biochemistry, the student will study proteins, carbohydrates, lipids, enzymes, and nucleic acids. Biochemistry has benefited greatly from the advancement of such techniques as chromatography (a technique for analyzing or separating mixtures of gases, liquids or dissolved substances), X-ray diffraction, and electron microscopy. These techniques have provided vast knowledge about how organisms obtain and store energy, how they manufacture and degrade their molecules, and how they sense and respond to their environment. Biochemistry is a vast field. We will scratch the surface of biochemistry in Unit Nine (Science 1109).

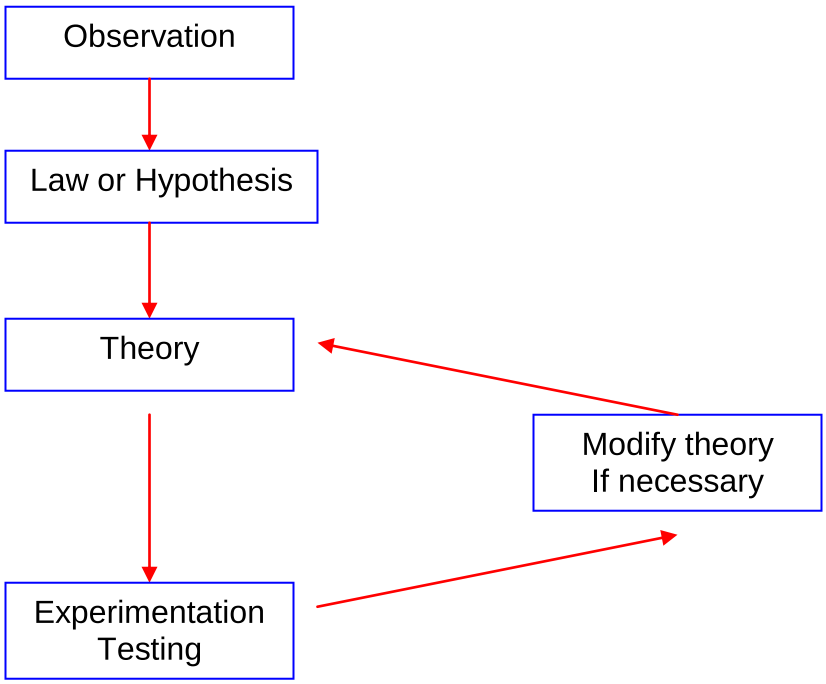
Pioneers in Chemistry:
Marie Sklodowska Curie. Madam Curie shared the Noble prize in 1903 with her husband Pierre Curie and Henri Becquerel for their work in radiation phenomena. In 1911, she won the noble prize for discovering radium and polonium. At that time, little was understood about the dangers of radiation. Much of her work was performed in inadequate laboratory conditions. She died in 1934 from the effects of exposure to radiation.

Figure 1.1

The Scientific Method

Many advances in science have come about by “accident”; however, they were not truly accidents because the people involved were using scientific methods of thinking. They were using logical steps of observation, deduction and experimentation to define something new to the scientific world. Most sciences utilize the process of the *scientific method.* The scientific method varies somewhat between the various fields of science such as psychology, zoology, anthropology, and the physical sciences. This course deals with the physical science aspects of the scientific method. The common steps are as follows:

1. **Observation** – The collection of data and the gathering of facts.
2. **Hypothesis** – Forming generalizations about the phenomena being studied. The finding of patterns, and the use of intuition.
3. **Theorizing** – Once all data and facts fit hypothesis, one can develop a working theory.
4. **Further testing**, experimentation, and evaluation. Back to step three if necessary.



Time to think:

1. What is good about chemistry?____
   ____
2. What is bad about chemistry?____
   ____
3. Formulate a hypothesis based on something recently observed.
   ____
   ____
4. How could this hypothesis be tested?
   ____

Teacher score numbers 3 and 4, initial here ____________________.

You are doing great! Now check your answers.

Number correct: ☐ Number incorrect: ☐

If you made any mistakes go back and make sure you understand the concepts.

Observing

Many valuable contributions to science came about because of people being very observant of their surroundings. Are you observant? Did you take a stairway today? How many steps were there? What is the current phase of the moon? How many utility poles are on your block? As Yogi Berra is credited to have said," You can observe a lot by watching". To be a good scientist, one must be aware of external events. Then, by recording the similarities and differences, one can see a pattern and develop a hypothesis. This hypothesis is then further tested and modified. As the hypothesis becomes accepted, it will become a theory. All theories are subject to more testing and experimentation. A theory cannot be proven. It is simply accepted until disproved. The scientific method works because of the cause and effect relationships in nature.

The "Black Box" laboratory Lab 1.1

One problem with chemistry experiments is that the scientist is often handling particles that are so small that they are invisible to the eye. Even powerful microscopes cannot detect an atom or a molecule. Thus, chemists must observe what happens and determine what is taking place at the molecular level without actually seeing the reactions, etc. One way to experience this phenomenon is to conduct the "Black Box" experiment. This lab calls for each student or group of students to be given a box (a shoebox will do) that contains a common household object. The students can use any means at their disposal to aid in hypothesizing what is contained in the box, without opening the box. The main idea is for the student to gather all of the *facts* that they can in order to form and test a hypothesis. Only when the students agree that they can no longer gather any more meaningful information that will help in the formation of a hypothesis can they open the box and test the hypothesis. Being right is not as important as being observant and gathering data intellectually.

Teacher score the Black Box Lab: ____________________________________

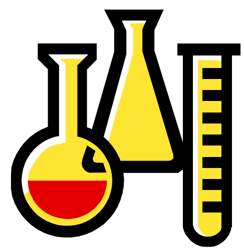
Natural Law, Theory and Hypothesis

Science is not always fact, but often theory. Some theories are good theories and therefore are accepted by most scientists. Other theories are weak with many scientists disagreeing about the validity or certainty of the theory. As new advances in technology and experimentation come along, theories are strengthened, adapted or discarded all together. A *natural law* is a descriptive principle that holds true in all circumstances covered by that law. If there are any exceptions under the circumstances stated by the law, then it cannot be considered a law. A natural law tells us what happens, whereas a *theory* attempts to tell us why it happens. For example, the law of gravity, which states that "every mass in the universe attracts every other mass with a force along a line joining them and that force is directly proportional to the product of their masses and inversely proportional to the square of the distance between them" (whew!), is a natural law. A theory for gravity would attempt to explain why this phenomenon is true, the presence of "gravitons" (theoretical gravity particles, not proven to exist), or something like that.

A *hypothesis* is a theory that has not become universally accepted and therefore suggests that it may not be true. For example, as new observations are made and new data is collected, old hypotheses are modified or discarded, and new theories may arise. As a result, science, as a collection of knowledge, is always dynamic (changing), never static (fixed).

The Language of Chemistry

Much of the confusion about the study of chemistry comes from the fact that many people assume that it is difficult. One of the difficulties may be that the language is so foreign to many students. A new vocabulary must be learned with new symbols and representations in order to understand chemistry. Do not be afraid of this new language. After a little practice, you will find that the language is a shortcut to expressing ideas and concepts. These shortcuts are utilized because scientists are a bit lazy about writing things out fully. For instance, it is a lot easier to write "Cl" instead of "Chlorine. Start using the symbols and abbreviations in your own note taking and the task of learning chemistry will become easier.



The data for physical science and chemistry in particular is in numerical form (numbers). These numbers are often plotted on graphs so that one criterion can be compared to another, such as time versus temperature. Science cannot proceed very far without some quantitative measurements. It is okay to say that the object is "heavy", but a scientist will want to know "how heavy?" The term "heavy" is" a relative term. For instance, a two hundred pound object is too heavy to lift, but ten times that amount would be classified as a "light" automobile.

Scientific Notation

Much of science deals with numbers that are either very large or very small. It is therefore necessary to find a means to express these amounts in a simplified way. *Scientific Notation* is a way of expressing very large or very small numbers. Scientific notation expresses a number as the number between 1 and 10, with the appropriate decimal, times a power of ten. To understand powers of ten, think about the fact that 100 can be written as 10 x 10, or 10^2. One thousand, 1,000, can be expressed as 10 x 10 x 10, or 10^3. We can also write 1,000 as 1×10^3. Thus, it can be seen that 1,500 can be written as 1.5×10^3. This shortcut greatly eases the trouble of expressing cumbersome numbers such as the speed of light, 300,000,000 meters per second, which can be written as 3.0×10^8 m/s.

Conversely, very small numbers can be expressed in negative powers of ten. For example, 0.005 grams of a substance can be written as 5×10^{-3} grams. It may appear to be just as easy to write 0.005, as it is to write 5×10^{-3}, but look at how cumbersome the mass of a proton would be to write out in long hand (1.673×10^{-24} grams = 0.000000000000000000000001673 grams!). See how the development of some scientific terms is due to the laziness of scientists? Go ahead, be lazy, and learn scientific notation.

In scientific notation, the standard is to place the decimal point after the first significant digit and adjust the exponent of ten so that there is no change in the value of the number. Think of the change as creating a new number with two parts, a digit part and an exponent part, from the old number. To change a number to scientific notation, put the decimal point behind the first digit (to the right of the first digit), divide or multiply the original number by the appropriate integer power of ten (just count the number of decimal places that you moved the decimal point). Then do the opposite (inverse) to the exponent part of the new expression so that there is no change in the value of the number.

For example, take the number 2,540. Move the decimal (which is currently to the right of the 0) to the right of the 2. Now since the decimal moved three places to the left, the power of ten will be to the third, or 10^3. Thus, the number expressed in scientific notation is 2.54×10^3.

On a scientific calculator, large numbers will be expressed as exponents. There may be seen an "E" on most calculators representing the exponent portion of the number. For example the number 2.78 million (2,780,000) will show up as 2.78 E6 on a calculator, which is the same as 2.78×10^6. This says to move the decimal 6 places to the right to get the actual number. For numbers less than one, the "E" portion of the figure will be negative, indicating to move the decimal to the left. For example five millionths (0.000005) may be displayed as 5.0 E-6, which is the same as 5.0×10^{-6}.

Time to think:

5. Express the number 1,000 in scientific notation.______________________

6. Express the number 0.001 in scientific notation.______________________

7. Express "one million" in scientific notation. ____________________Enter one billion into a scientific calculator, press "=" or "enter" and see what happens to the number. Write the number that appears on the calculator here______________________.

8. Write 1,206 in scientific notation.___________________________________

You are doing great! Now check your answers.

Number correct: ☐ Number incorrect: ☐

If you made any mistakes go back and make sure you understand the concepts.

Units

The *units* of a measurement tell us the scale, standard or type of object that is being measured. For example, we may have 4 cars, 12 kilograms, or 5 centimeters. The numbers: 4, 12, and 5 tells us how much and the *units*: cars, kilograms, and centimeters tell us *what* is being measured or tallied. This is critical. If I ask my employer how much I will be paid, she might say “four”. Is this $4,000 a month, $4.00 an hour, or what?

The most common standards for measurement are the “*English standard*”, which is used in the United States, and the “*metric system*”, which is used throughout most of the rest of the world. Scientists prefer the metric system because of the fact that it is much easier to convert very large or very small quantities when using the standardized metric system. The *Systeme Internationale* (SI) was established in 1960. It is an international agreement as to the system of units that would be used in the scientific community. It is based on the metric system and is comprehensive in its scope (which means it contains all that is needed). The SI fundamental units (table 1.1) can be manipulated to form almost every other unit of measurement. For instance, velocity is m/s (meters per second), area is cm^2 or m^2, volume is cm^3 or m^3, and density is usually in g/cm^3 (grams per cubic centimeter).

Physical Quantity	Unit	Abbreviation
Mass	Gram	g
Length	Meter	m
Temperature	Degree (Kelvin)	K^o
Time	Second	S
Charge	Coulomb	C
Energy	Joule	J

Table 1.1 The Fundamental units

The conversion of the metric units often involves *prefixes* to denote various factors of ten. It is much easier to convert large units to small units in the metric system after one is familiar with the powers of ten. For example, it is much easier to figure how many centimeters are in a kilometer (100 cm per meter and 1,000 m per kilometer = number of cm per kilometer) than it is to figure how many inches are in a mile (12 inches per foot times 5,280 feet per mile = number of inches per mile). The math is much easier for conversions in metric. The prefixes for the metric system are listed in table 1.2.

Prefix	Symbol	Value	Example	Example name
Peta	P	10^{15}	Ps	Petasecond
Tera	T	10^{12}	Tm	Terameter
Giga	G	10^{9}	GW	Gigawatt
Mega	M	10^{6}	MC	Megacoulomb
kilo	k	10^{3}	km, kg	kilometer, kilogram
centi	c	10^{-2}	cm	centimeter
milli	m	10^{-3}	mg	milligram
micro	μ	10^{-6}	μ m	micrometer
nano	n	10^{-9}	ns	nanosecond

Table 1.2 Metric prefixes

When manipulating numbers, such as in formulas, it is crucial to maintain the proper units. This process is sometimes called dimensional analysis. In dimensional analysis, we begin with the “known quantity”. (The known quantity is the number and units of material that are given at the start of the computation. In the example below, equation 1.1, the known quantity is 3.2 lbs). Place the known quantity in the numerator of the beginning fraction if there is no denominator. To use dimensional analysis, one must know the dimensions of units and conversion factors (see table 1.3). If we want to find the number of grams in 3.2 pounds of salt, we begin with:

$$3.2\ \cancel{\text{Pounds}} \times \frac{16\ \cancel{\text{ounces}}}{\cancel{\text{Pound}}} \times \frac{28.35\ \text{grams}}{\cancel{\text{Ounce}}} = 1451.52\ \text{grams} \quad \text{(Eq.1.1)}$$

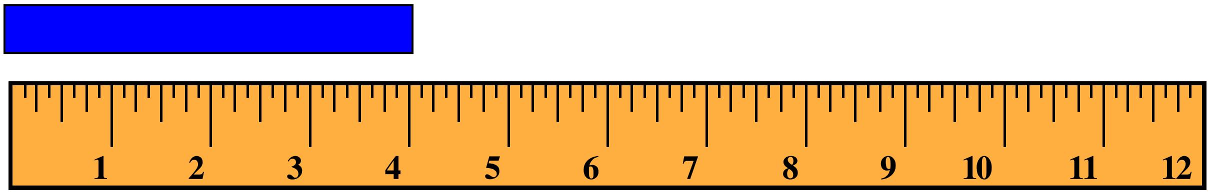
Notice that the units in the denominator divide through by one (cancel out) the same numerator units, anything divided by its self equals one. In the above equation, the pounds and ounces cancel out and all that is left is the grams. It is very important to see where we want to go in the conversion. Just as it could be written that one pound is 16 ounces, 16 ounces is one pound: 1 pound / 16 ounces or 16 ounces / 1 pound. To determine which way to write the fraction, keep in mind that we want one unit in the denominator and the same unit in the numerator in order to cancel (divide to get one).

SAE	Metric
1 inch	2.54 cm
1 mile	1.609 km
1 gallon	3.75×10^{-3} m^3
1 atmosphere	1.013×10^5 Pascal
1 KWh	3.6×10^6 Joule
1 ounce	28.35 grams
1 quart (liq.)	0.94635 liter

Table 1.3 Selected conversion units

Significant Figures

Measurements can never be more accurate than the device used to measure the quantity. A ruler or meter stick is fine when measuring lengths for the cutting of lumber or figuring the amount of tile required for the kitchen floor, but if one were measuring the diameter of a bearing for a water pump, one would need to make the measurements with a micrometer or set of calipers. In figure 1.2, the length of the solid line can be seen as approximately 4 inches. When stated as thus, it is assumed that the measurement is accurate to within plus or minus (+/-), 1 inch. Any further speculation into the length of the solid line is “estimation”. One may be able to estimate the length as 4 1/8 inch. (An accuracy of +/- 1/8 in.), but any further accuracy is indeterminate due to the lack of any finer increments of gradation on the ruler. When writing “4 inches”, the 4 is assumed the *uncertain* part of the figure. When writing “4 1/8 inches”, the 1/8 is the uncertain portion of the figure.



(in.)

Figure 1.2

The numbers that arise from measurements are subject to *significant figures*. When a number is written, it is assumed that the final digit, or rightmost digit, is uncertain. This helps us to determine the degree of accuracy of the measurement. When performing an arithmetical computation of figures, such as adding, the answer can only be determined to have a precision that is the same or less precise than the least significant figure.

Thus, in the sum:

$$\begin{array}{r} 4.125 \\ \underline{200.8} \\ 204.925 \end{array}$$

We must consider that the least accurate significant figure is the “0.8” in the number 200.8. Therefore, the sum cannot reflect accuracy any finer than to the tenth. The above sum must be *rounded off* to the nearest tenth, or 204.9.

Time to think:

9-11. What do you think would be the most reasonable metric units to use to measure the:

9. Distance from San Diego, California to San Francisco, California? ____

10. The length of a Semi-truck? ________________________

11. The thickness of a piece of paper? ____________________________.

12-14. Using the information from table 1.3, calculate:

12. The number of ounces in a gram _____________.

13. The number of miles in a kilometer _________________________

14. The number of quarts in a liter __________________________________

You are doing great! Now check your answers.

Number correct: ☐ Number incorrect: ☐

If you made any mistakes go back and make sure you understand the concepts.

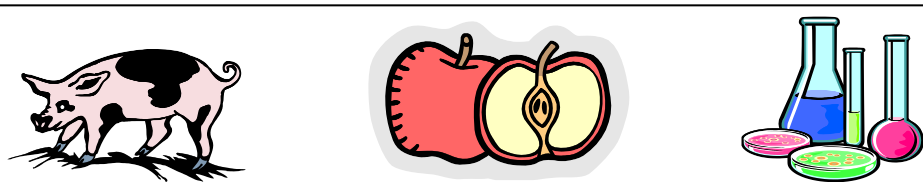
Precision and Accuracy

The measurement of any physical quantity is subject to some uncertainty. The *precision* of a measurement refers to the agreement of two or more measurements of a value. *Accuracy* is the agreement between the measurement and the accepted or true value. As an example, a student may measure the mass of a magnesium ribbon four times and find that her values differ from each other by no more than 1%. Let us say she measures the magnesium ribbon on a triple beam balance and records weights of 4.044 g, 4.039 g, 4.047 g and 4.046 g. All of these measurements are very close to each other. This is fairly *precise*. However, if the *known value* of the magnesium ribbon is 10% more than her average measured value (say 4.400 g), then her measurements were not *accurate*. The difference may lie in her method of measuring or in the equipment that is used; either way, it cannot be considered an accurate measurement.

Matter

Whenever students are asked, "What is matter?" inevitably, the word "stuff" is used. Matter is indeed the "stuff" of which the universe is made. *Matter* is defined as "that which occupies space and has mass". Think about this. Almost everything that we see and interact with is matter. In fact if it is not matter, it is either pure energy (like light) or abstract thought or emotions. At this point, differentiation should be made between mass and weight. *Mass* is defined as; "the property of an object that is a measure of its *inertia* (the property of a body by which it remains at rest or continues moving in a straight line unless acted upon by a directional force). Mass is simply "the amount of matter an object contains". *Weight* is the force upon something that is caused by gravity. All this really means is that mass is the amount of matter that we have, whereas weight is dependent upon gravity. The mass of something is always the same, but the weight will vary with the effects of gravity. On the moon a person would weigh 1/6 as much as they do on the earth, because the pull of gravity is 1/6 that of the earth, but the mass will be the same any where in the universe. On the surface of the earth hydrogen gas rises, it has a negative weight. The pull of gravity is not strong enough to hold such a small mass. Yet there can be no such thing as negative mass. The mass of hydrogen is the same anywhere in the universe (1.008 g per mole). Mass is how much of the matter there is, weight is how the matter is affected by gravity as measured on a scale.

Matter is divided into four physical states of matter: solid, liquid, gas and plasma. A *solid* is a substance that when acted upon by a force tends to maintain both volume and shape. True solids are crystalline, that is they have a regular polyhedral structure to them. An amorphous solid is a solid with no regular shape to it. Glass is an example of an amorphous solid. A *liquid* has a definite volume but takes the shape of its container. The liquid state of a substance is usually found in the temperature range between that of a solid and that of a gas. A good example is



Anything that has mass and takes up space is called matter.
Fig. 1.3

water. Water is in the liquid state in the temperature range between ice (0°C) and steam (100°C).

A *gas* has no fixed volume or shape. A gas will evenly fill the volume and shape of its container. Without a container, it will tend to expand without limit. A fourth state of matter, called *plasma*, exists when a gas becomes ionized. Plasma exists inside stars and in interstellar gases. Just think of plasma as super hot goo, a soup of electrons not associated with any nucleus and positively charged nuclei randomly bouncing around at very high speeds. On Earth, we find matter only existing in the three physical states, solids, liquids and gases (except for plasma found in lightning strikes).

Matter can also change to energy and vice-versa according to Einstein's famous equation:

$$E = mc^2 \quad \text{eq.1.2}$$

Where E is the energy, m is the mass and c is the speed of light. Matter can undergo changes in physical states. However, matter and energy cannot be created or destroyed they can only be transformed. This is the first law of thermodynamics or the law of conservation of energy. Except in a a nuclear reaction, the masses (amount of matter) of reactants will always be equal to the masses of the products in any chemical reaction.

Solids

In a *solid*, the atoms are held together tightly. They do not move around that much. They just vibrate as temperatures go up. Thus, solids are rigid and they do not compress easily. In ionic solids such as table salt crystals, the ions are connected to their neighbors by electrical attraction. Crystals such as diamonds produce the hardest materials. Ice is an excellent example of a crystalline solid. The H_2O molecules are held together in a rigid structure. In fact, for ice to form the crystal, the molecules must back away from each other slightly to obtain the perfect bond distance. This makes water at 0^oC less dense than water above 5^oC. (This is a very good thing, because if ice were denser than liquid water, the oceans would freeze from the bottom and completely ice-over, due to the ice at the bottom of the oceans never getting enough warmth from the sun to melt. Eventually, the ice would expand up to the surface and remain as ice).

As heat is added to a solid, the molecules vibrate more and more until a certain temperature is reached where the molecules break their hold on each other and are free to move about. This certain temperature is called the *melting point* (m.p.). In reverse, it is known as the *freezing point*. Most solids have a melting point, even rocks. Lava and magma are molten rock. There are a few solids that go directly from the solid state to the gas state without passing through the liquid state; carbon dioxide at atmospheric pressure is an example. Other solids, such as sucrose or "table sugar", decompose rather than melt before reaching the liquid or gaseous state.

Liquids take the form of their container

Liquids

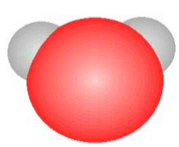
Liquids are materials in which the atoms or molecules are approximately as close to each other as solids, but the materials can slip and slide over and around each other to change places. This is why a liquid will take the form of the vessel in which it is contained. The positions of the molecules are no longer fixed, but since they are still very close to each other, the volume of the liquid is about the same as the volume of the solid.

Gases

Matter that exists in the *gas* phase exists in the most energetic form. As a liquid is heated, the molecules get more excited and vibrate more. At a certain temperature, under certain pressure, the molecules will escape the attraction for other like molecules, and be hurled into the atmosphere. This certain temperature is called the *boiling point* (b.p.). These "escaped" molecules now exist in the form of a gas. The individual molecules are not bound to each other and will continue to move in the direction that they are traveling until acted upon by a force, such as the walls of the vessel that contains the gas.

One method of visualizing the three phases is to imagine the balls of a billiard table. Let each of the balls represent a molecule. In the solid form, the balls are racked up and unable to move. Even when the rack is removed, the solid retains its shape. Just as an ice cube will retain its shape after being removed from the ice cube tray. The molecules have a definite orderly design and maintain that geometric shape as if they were glued to each other. In the liquid form, the billiard balls are free to move about each other, but they will stay very close to each other. It is as if someone tipped the pool table at an angle and all the balls would collect at the low end of the table. The balls can exchange locations with other balls, but they would all stay close enough to touch one another. In the gas phase, it is as if the rack of balls has just been "broken" and every ball is moving in a different direction. With the exception that each ball is not hindered by gravity or friction, the balls keep moving about, bouncing off the rails continually. If there were, no rails the balls would sail off into space and not change direction until acted upon by some other force. Now imagine this billiard table analogy in three dimensions to picture the three states of matter.

In day-to-day life, we are constantly dealing with solids, liquids and gases. When we pour our orange juice for morning breakfast, we are careful to use a cup or glass so as not to spill it. One cannot spill a solid, such as an orange or toast. If we do spill the juice, we see that it quickly goes all over the place. It takes the form of the vessel in which it is contained. A gas, such as the steam escaping from the teakettle, will fill the room. Gases must be contained. It is easy to see that when heated, the water in the teakettle reaches a point where the H_2O molecules loose their hold on each other and zing their way into the atmosphere. As the molecule cools, it condenses and collects as liquid water.



Water molecule

Time to Think:

15. Name three materials that are solid? ______________________________________

16. Name three materials that are liquid? ______________________________________

17. Name three materials that are gas? ______________________________________

Have your teacher initial below for questions 15-17.
You are doing great! Now check your answers.

Number correct: [ ] Number incorrect: [ ]

If you made any mistakes go back and make sure you understand the concepts.

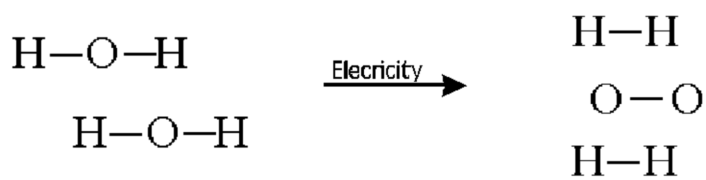
Way to go!!!

Mixtures

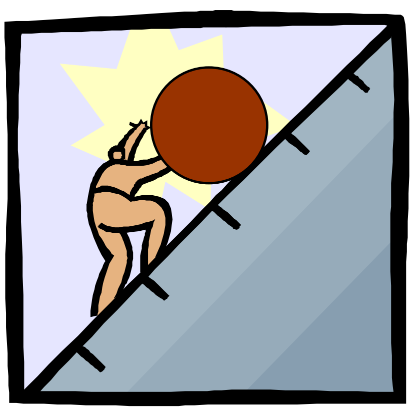
Some substances are difficult to distinguish. In the real world, most substances are not pure. They are a mixture of solids, liquids and gases. Bread has air pockets, which makes it a mixture. Peanut butter has the right combination of fats and solids to give it a texture somewhere between a solid and a liquid. Pure substances are substances that have constant composition. Pure water always behaves the same at the same pressure and temperature. *Mixtures* have two or more substances, which are variable in composition. Mixtures and the separation of mixtures will be covered in chapter 7.

Compounds

Compounds differ from mixtures in that the combination of elements and/or other compounds involves a chemical reaction. A *physical change* does not effect the composition of a substance. For instance, the freezing of water or turning water into steam does not affect the water chemically. Likewise, mixing sand with water does not affect either substance chemically (another example of a physical change). However, if the right amount of electrical current is run through water, the H_2O is broken into its component parts and we will have hydrogen and oxygen gas. This is a chemical reaction, resulting in twice as much hydrogen gas as oxygen gas.



Decomposition of water Fig. 1.5



Pushing the ball up the hill adds potential energy to the system. Fig. 1.6

Energy

Energy is a measure of a system's ability to do work. When something possesses energy, it can affect other objects by performing work on them. There are two major kinds of energy, *potential energy* and *kinetic energy*. Potential energy is energy that is a result of an object's position or composition. The kinetic energy of an object is due to motion of the object and depends on the mass and velocity of the object. The following equation determines the kinetic energy of an object, when given its mass and velocity:

$$KE = \frac{1}{2} mv^2 \qquad \text{eq. 1.3}$$

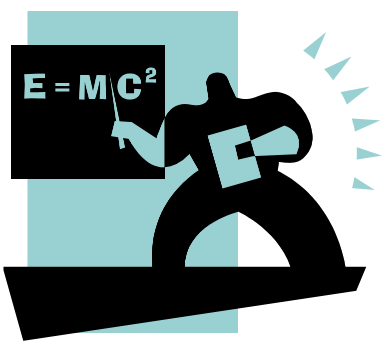
Where KE is the kinetic energy, m is the mass of the object in kg, and v is the velocity of the object in m/s (meters per second). There is no need to memorize the above equation; you will see it again in physics class.

If an object falls or rolls downhill, it gains kinetic energy as it accelerates. From where did the energy come? Potential energy is energy associated with height in this instance. Work was done to raise the object, to overcome the pull of gravity. Now the gravity can be used to do work. The potential energy is "stored" in it. Falling or "downhill moving" objects convert potential energy into kinetic energy, but the total energy stays the same. Remembering that energy is the ability to do work, we could also say that potential energy is stored work.

In chemical reactions, there is usually energy given off (an *exothermic* reaction) or energy required (an *endothermic* reaction). The energy given off is from the energy that was stored in the chemical bonds. It took work (energy) to form the chemical bonds in many substances, and when those bonds are broken, energy is given off, usually in the form of heat. This is what is taking place when gasoline is burned in an internal combustion engine or when carbohydrates are converted into energy for the human body. Chemical bonds are broken and heat is given off.

In the same way, it may require energy to form chemical compounds. An example of this is the sunlight energy required to form carbohydrates from CO_2 and O_2, such as green plants accomplish through photosynthesis (see Unit Ten). Thus, potential energy is stored in the chemical bonds of the plant's components. The sugars and starches in the plant have stored up energy that can be utilized later. We see this energy released when animals ingest the plants to produce food calories, or when wood or gasoline is burned resulting in heat and light.

Calories and Specific Heat



The units for energy will depend on the system being measured. As stated earlier, in chemical reactions, the energy is usually released or absorbed as heat. The unit for heat is the *calorie*. A calorie is defined as "the quantity of heat necessary to raise the temperature of 1 gram of water 1 ^{o}C at 1 atmosphere pressure and 15^{o}C". Therefore, if a gram of water is to be heated from 14^{o}C to 15^{o}C, it will require 1 calorie of heat (energy). A kilocalorie is 1,000 calories. The kilocalorie is equivalent to one food Calorie. The calorie is abbreviated cal and is distinguished from the food Calories, which should always be written with an upper case C. One calorie is equal to 4.180 Joules (refer to table 1.1). The Calorie is obsolescent (out of date) and the calorie is being replaced by the joule (one calorie = 4.1868 joules).

Another property of matter is the *specific heat capacity* or *specific heat*. This is the amount of heat required to raise the temperature of 1 g of a substance 1^{o}C. This is important because different substances respond to the addition of heat differently. For instance, liquid water has a heat capacity of 4.184 J/g ^{o}C (Joules per gram, 1 degree Celsius). Whereas water in the form of ice has a heat capacity of 2.03 J/g ^{o}C and gold 0.13 J/g ^{o}C! This means that water can absorb more energy, in the form of heat without increasing temperature, than gold. This is what makes gold feel cooler to the touch than water. The gold takes away more body heat.

We can figure the energy required to heat a known quantity of a known substance by the use of the formula:

$$Q = s \times m \times \Delta T \qquad \text{eq. } 1.4$$

Where

Q = Energy (heat), in joules required
s = specific heat capacity
m = mass of sample (g)
ΔT = change in temperature (^oC).

Chapter One Review

Teacher's Initials ____________

(Table 1.3 may be used on all reviews and tests in this unit.)

SAE	Metric
1 inch	2.54 cm
1 mile	1.609 km
1 gallon	$3.75 \times 10^{-3}\ m^3$
1 atmosphere	1.013×10^5 Pascal
1 KWh	3.6×10^6 Joule
1 ounce	28.35 grams
1 quart (liq.)	0.94635 liter

Table 1.3 Selected conversion unit

1. Why is chemistry an important topic?

____

2. Why is observation necessary to the scientific method?

____

3. If experimental data does not support the hypothesis, what choice(s) does the scientist have?

____

4. How does "natural law" differ from "theory"?

____

5. Express the following in scientific notation:
   a. 101,000 ________________________________
   b. 6,600,000 ________________________________
   c. 0.000046 ________________________________
   d. 0.0000000020 ____________________________

6. How many centimeters are there in 2.2 miles? ________________

7. How many grams are in a pound? __________________________

8. A megawatt is how many watts? (Answer in scientific notation form).

____

9. What is the difference between precision and accuracy?

____

10. The pull of gravity on the moon is 1/6 what it is on the earth. The mass of a 120 kg man on earth would be ___________kg on the moon.

11. Give an example of water existing as three states of matter.

____

12. Why don't liquids compress easily?

____

13. When does a baseball bat have more kinetic energy, just before or just after hitting a pitched fastball? Explain.

____

14. The specific heat capacity for aluminum is 0.89 J/g ^{o}C. How much energy (in Joules) is required to heat 12 grams of aluminum from 21 ^{o}C to 52 ^{o}C?

____

15. In the same classroom at constant temperature, why do you think metal objects feel cooler than wooden objects?

____

You are doing great! Now check your answers.

Number correct: ☐ Number incorrect: ☐

If you made any mistakes go back and make sure you understand the concepts.

Peaceful

"Nothing can bring you peace but yourself; nothing, but the triumph of principles"

-Ralph Waldo Emerson

Peaceful

Chapter Two

Character Quality: Peaceful

"*Peace cannot be achieved through violence; it can only be attained through understanding.*"
Albert Einstein (1879-1955)

Seek peace; allow inward tranquility to control your heart and mind. You are peaceful with others to the extent you are peaceful with yourself, because a stable mind is a peaceful mind. Invite peace into your life and embrace it. Make it the very state of your mind. Peace does not stay with you unless you develop certain qualities and establish certain conditions in your life. Without self-discipline, love, control of your thoughts and habits and desires, joy, and appropriate attitudes, you cannot expect to have peace in your life.

"Work joyfully and *peacefully*, knowing that right thoughts and right efforts inevitably bring about right results." *James Allen*

The Atom

What is the smallest size that an object can become? How do we know that things exist that we cannot even see? Is the atomic theory just an idea that someone made up to fool all of us? Twenty-six centuries ago one of the most remarkable men was born. He lived in an area that was then known as Greece. His name was Thales. He is considered the founder of mathematical reasoning. He was the first Greek *philosopher* (meaning "lover of wisdom"). Thales studied math, astronomy, agriculture and the elements. He decided that all matter was made up of different forms of the element water. He was wrong, but the concept of everything being composed of tiny particles launched the science of alchemy and eventually chemistry. "Chem" was the Egyptian word for black (from "kema" for "soil"), and many people think that the word "chemia", which became "chemistry" came from this Egyptian word (at the time, some thought of alchemy as "black" magic).

In the 5th century B.C., the Greek philosophers Democritus and Leucippus proposed that matter was made up of tiny, indivisible particles in constant motion. They called these tiny particles "atoms" from the Greek word "atomos", spelled "ατθμθσ" in Greek. It means "indivisible". Democritus and Leucippus are known as the "atomists". There are five major points to their atomic idea:

1. - All matter is composed of atoms, which are bits of matter too small to be seen. These atoms cannot be further split into smaller portions.
2. - There is a void, which is empty space between atoms.
3. - Atoms are completely solid.
4. - Atoms are homogeneous, with no internal structure.
5 - Atoms are different in their size, shape, and weight.

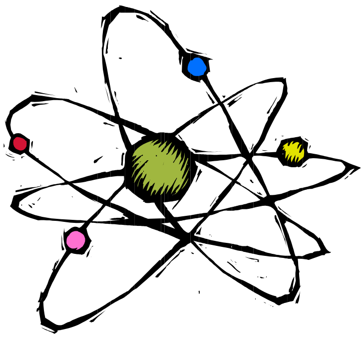


Figure 1.7 The Atom

Aristotle, the famous philosopher, however, did not accept the atomist's theory, and it was ignored for centuries. In 430 B.C. Greek philosopher Empedocles proposed that all matter is made up of four elements; air, water, fire, and earth. Modern atomic theory began with the publication in 1808 by John Dalton of his experimental conclusions that all atoms of an element have same size and weight. Dalton further proposed that atoms of elements unite chemically in simple numerical ratios (as molecules) to form compounds. The precision chemical balance of that day enabled accurate measurements of compounds and led to the detection of several elements. In 1911 Ernest Rutherford explained an atom's structure in terms of a positively charged nucleus surrounded by negatively charged electrons (which were not discovered until 1899) orbiting around it. In 1913, Niels Bohr used quantum theory to explain why electrons could remain in certain allowed orbits without radiating energy. The development of quantum mechanics during the 1920s resulted in a satisfactory explanation of all phenomena related to the role of electrons in atoms.

The *atom* today is defined as "the smallest part of an element that can exist chemically". This is not much different from the definition given by the atomists. Hmmm, so what is an element? An *element* is a substance that cannot be decomposed into simpler substances. In an element, all the atoms have the same number of protons in the nucleus. Each element has unique properties.

There are 92 naturally occurring elements. Atoms consist of very small and dense nuclei (plural for nucleus) consisting of protons (positively charged) and neutrons (neutral, no charge) surrounded by rapidly moving electrons (negative charge). In a neutral atom, the number of electrons equals the number of protons. The number of neutrons can vary. Atoms can bond with other atoms to form molecules. Atoms are the building blocks for molecules. A *molecule* is two or more chemically bonded atoms. A *molecule* is the smallest part of a compound having all the properties of the compound.

Imagine being unaware of the atomic theory. As you cut through something, say a piece of wood, it may be cut in half, half again, and again. At some point you may ask, "What is the smallest size that I can cut this wood and still have it be wood?" This is what the atomists asked. If we cut a long piece of cotton rope in half, we still have a piece of rope. At what point of "halving" does the rope not be a rope anymore? When we cannot tie it into a knot? At what point is it no longer cotton? If we keep cutting anything in half, eventually we will reach the molecule. Cutting this molecule will change the properties of the substance and so we no longer have the same substance.

Cutting a water molecule (H_2O) into its constituent parts of hydrogen and oxygen will change the properties and we will no longer have water. Hydrogen and oxygen behave completely different from water. Cutting the hydrogen or the oxygen into pieces cannot be done by any common means of separating matter. If done, the "splitting" of the atom will release the subatomic particles and a great deal of energy. This would be an atomic reaction.

Time to Think:

1. Define “atom”____
2. In your opinion, why is the atomic model important?
3. What is an “element? ____________________________________

You are doing a good job and having fun, right? Check your answers.

Number correct: ☐ Teacher initial number 2 ☐

If you have any errors go back and make sure you understand the concepts.

John Dalton

Pioneers in chemistry: John Dalton

All atoms share the same basic structure. The model for the atom has undergone some change over the past 200 years. In the early 1800’s, John Dalton established the first useful atomic theory of matter. As a child, Dalton attended the Quaker school at Pardshow Hall in Eaglesfield, England, at a time when only one out of every 215 English people could read. John was quick at studies and tireless at mathematical problems. In 1778, when John was twelve, he opened a school in Eaglesfield, England.

As a young adult Dalton worked at self-improvement, including answering questions from women’s and men’s magazines. His responses appeared in print sixty times. He found a friend and mentor in John Gough, the blind son of a wealthy tradesman. Gough taught Dalton languages, mathematics and optics, and shared his extensive library with him. As Dalton's interest in science expanded to include optics, pneumatics, astronomy and geography, he began in his early 20’s to supplement his low income with public lectures. He also approached a nearby museum with an offer to sell his eleven volume classified botanical collection. He collected butterflies and studied snails, mites and maggots by suspending them in water and in vacuums.

At the age of 27, John Dalton moved to Manchester and became a tutor at New College. As a chemistry tutor, Dalton taught from Lavoisier's *Elements of Chemistry*. That same year he published his first book, *Meteorological Observations and Essays*. In it, he said that each gas exists and acts independently and purely physically, rather than chemically. This means that gases act according to mechanical repulsion rather than chemical attraction. His interest in gases arose from his meteorological studies. Dalton always carried his weather apparatus with him wherever he went. He was constantly studying the weather and atmosphere. During his lifetime, he recorded over 200,000 observations in a journal, which was his constant companion. It was in these observations that his mathematical mind saw the numerical connections between the data that he had collected.

At 37, while attempting to explain his law of partial pressures, Dalton started to formulate his most important contribution to science, the *atomic theory*. He was studying nitrogen oxides for Doctor Priestley's test for percentage of nitrogen in the air. Among the reactions, he studied were those of nitric oxide with oxygen. He discovered that the reaction could take place in two different proportions in exact ratios, namely:

$$2NO + O \rightarrow N_2O_3$$
$$NO + O \rightarrow NO_2 \qquad \text{eq. } 1.5$$

Dalton stated that oxygen combines with nitrogen sometimes 1 to 1.7 and at other times 1 to 3.4 by weight. Later that same year, he stated the law of multiple proportions, which states that the weights of elements always combine with each other in small whole number ratios. Dalton published his first list of atomic weights and symbols that year, thus giving chemistry a language of its own.

John Dalton's atomic model contained four main concepts:

1. All matter is composed of tiny indivisible particles called atoms.
2. All atoms of each element are exactly the same.
3. Atoms of different elements have different masses.
4. Atoms of different elements can join to form molecules or compounds.

Not much different from the five points proposed by the atomists Democritus and Leucippus!

"Could anything at first sight seem more impractical than a body which is so small that it's mass is an insignificant fraction of the mass of an atom of hydrogen?"
-- J.J. Thomson.

"Every kind of *peaceful* cooperation among men is primarily based on mutual trust and only secondarily on institutions such as courts of justice and police"
-Albert Einstein

Joseph John Thomson

The Thomson Model of the Atom

Almost 100 years after Dalton at the Cavendish Laboratory at Cambridge University, J.J. Thomson was experimenting with currents of electricity inside empty glass tubes. He was investigating a long-standing puzzle known as "cathode rays". When a glass vacuum tube was connected to a source of high electricity, glowing green "rays" appeared. Thomson discovered that these rays were bent by electricity, so he hypothesized that they must be charged particles. His experiments prompted him to propose a hypothesis: "These mysterious rays are streams of particles much smaller than atoms; they are in fact minuscule pieces of atoms. Since atoms are electrically neutral, whatever particle was charged had to have an oppositely charged particle to balance it". At first, Thomson called these particles "corpuscles" later he named them "electrons" (the particle that he is credited with discovering in 1899). Thomson imagined the atom as a "chocolate chip cookie dough" sphere with tiny electrons imbedded like the chocolate chips. The "dough" is positively charged to keep the atom neutral. It is called the "plum pudding" model of the atom.

> "The general public may be able to follow the details of scientific research to only a modest degree, but it can register at least one great and important gain: confidence that human thought is dependable and natural law is universal."
> --Albert Einstein

In 1911, Thomson's own former student, Ernest Rutherford, struck down the plum pudding theory. Using a different kind of particle beam, Rutherford found evidence that the atom has a small core, a nucleus. Rutherford suggested that the atom might resemble a tiny solar system, with a massive, positively charged center circled by only a few electrons. It was found that this nucleus was built of new kinds of particles (protons and neutrons), much heavier than electrons. Rutherford conducted the now famous gold foil experiment in which he aimed a beam of positively charged particles (*protons* that he was credited with having discovered in 1919) at a very thin piece of gold foil, only a few atoms thick. Most of the particles passed straight through, but some bounced back showing that the particles had struck something larger than a single proton. He called this region the *nucleus*. Rutherford's model of the atom presented a nucleus with positively charged protons and almost all of the atom's mass, surrounded by mostly empty space and circled by electrons.

Bohr's Model

Two years later Niels Bohr modified Rutherford's model of the atom. Bohr was born and educated in Copenhagen, Denmark. After receiving his doctorate 1911, Bohr went to England to study with Sir J.J. Thomson at Cambridge. He had intended to spend his entire study period in Cambridge but did not get along well with Thomson. After a meeting with Ernest Rutherford in Cambridge in December of 1911, Bohr moved to Manchester in 1912. There he worked with Rutherford's group on the structure of the atom. Rutherford became Bohr's role model both for his personal and scientific qualities. In 1913, Bohr proposed that each electron in an atom has a fixed amount of energy. This energy keeps the electron moving about in an orbit around the nucleus at a fixed energy level. These energy levels are like rings or shells that build up around the atom in larger and larger shells, like the layers of an onion. This theory supported the new idea of quanta put forth by Max Planck just 12 years earlier. If the atom absorbed energy, the electron jumped to a level further from the nucleus; if it radiated energy, it fell to a level closer to the nucleus. Bohr's model, called the "*planetary model*", or sometimes

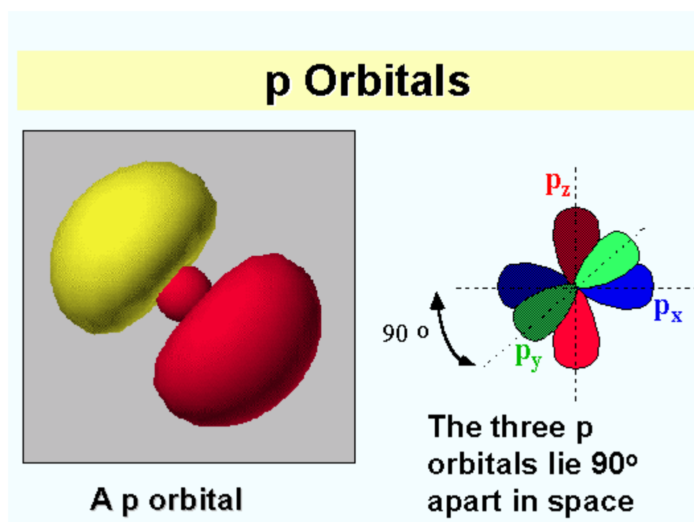


Figure 1.8

called “Bohr’s Theory”, was a huge leap forward in making theory fit the experimental evidence that other physicists had found over the years. He received a Nobel Prize for this work, as well as for his work on radiation, in 1922.

The atomic model has since been further modified to what is now called the electron cloud model. Today scientists know that electrons do not travel in actual orbits around the nucleus. In current models of the atom, electrons form clouds, which travel in spherical, and other shaped orbitals around the nuclei. Those closest to the nucleus have a spherical orbital called the 1s (pronounced “one es”) orbital. The 1s orbital contains up to two electrons. In the second energy level, the first two electrons also travel in a larger spherical orbital called the 2s (“two S”) orbital. Up to six more electrons occupy dumbbell-shaped 2p orbitals, each of which contains two electrons. The three 2p orbitals are arranged at right angles to each other, centered along the x, y, and z-axis (fig. 1.8). Additional electrons form the more complicated d and f orbitals. This topic will be explored in further detail in later chapters.

The electron clouds are not so difficult to visualize. A fan blade moving very fast appears as a blur. Likewise, electrons move so fast that they cannot be said to occupy any given place at a given time. The paths of the electron around a nucleus take up almost all of the atom’s volume. To get some sense of proportion, if an electron were the size of a softball (approx. 10 cm diameter); the nucleus of an average size atom would be almost 1 Km in diameter. If this size of a nucleus were at the center of the earth, the sphere of an average atom’s electron cloud would almost reach the surface of the earth. Using this comparison, it is easy to see that an atom is mostly “empty space”. (It may be of some comfort the next time you hit your head against a brick wall, to know that the wall is mostly “empty space”).

Time to Think:

4. In Rutherford’s Gold foil experiment, what was it that prevented some of the protons from passing through the foil? ______________________________
____.

5. What is another name for “Bohr’s Theory” of the atom?
____

6. Why do scientists state that an atom is mostly “empty space”?
____

You are doing well! Now check your answers.

Number correct: [ ] Number incorrect: [ ]

If you made any errors, go back and review the concepts.

Chemical Bonds

A pencil "lead" contains pure carbon

When atoms combine chemically, they create a chemical bond. A *chemical bond* is an attractive force that holds atoms or ions together. This attraction may be thought of as atoms holding each other's hands. The manner in which atoms are bound together has a profound effect on the chemical and physical properties of the substances. For example, graphite, diamond and coal are composed solely of carbon atoms. Coal is a black soft substance thought to be formed from the remains of ancient organic matter. Graphite is a soft, slippery, dark gray material that is used as a lubricant and in pencil "lead'. Diamond is a precious gemstone that is one of the hardest materials known to man. Diamonds are used in jewelry and in industrial cutting tools. Why is it that these three materials, which are composed solely of carbon, can have such different physical properties? The answer lies in the different ways in which the carbon atoms are bonded to each other in the substances. We will further explore the way carbon bonds in the unit on organic chemistry.

The following is a series of chemical formulas for some well-known chemical compounds:

HCl: hydrochloric acid;
H_2O: Water
NH_3: ammonia
CH_4: methane

A diamond is pure carbon

Notice that each formula contains hydrogen. The first formula shows that one atom of hydrogen (H) combines with one atom of chlorine (Cl) → (HCl). The second formula shows that two atoms of hydrogen combine with one atom of oxygen (O) → (H_2O). The third formula shows that three atoms of hydrogen combine with one atom of nitrogen (N) → (NH_3). The fourth formula shows that four atoms of hydrogen combine with one atom of carbon (C) →(CH_4). Why do different amounts of hydrogen atoms bond with these different elements? In each case, the hydrogen atom is the same. Therefore, the element to which the hydrogen atom is bonded must be different in order to require varying amounts of hydrogen. To understand the difference, we must remember that an atom is made up of a small, positively charged nucleus surrounded by a cloud of negatively charged electrons. The electrons are involved in the bonding. We will learn more about chemical bonding and the role of the outer electrons in units two and three.

Thermodynamics

Thermodynamics is the study of the laws that govern the conversion of energy from one form to another, the direction in which heat will flow, and the availability of energy to do work. It is based on the concept that any system will have a certain amount of internal energy. This internal energy is the total kinetic energy and potential energy of the atoms and molecules of the system that can be transferred directly as heat. In chemistry, the study of heat is very important because as a rule, as systems heat up, the molecules move faster, and react quicker. We will learn more about thermodynamics in Unit Four.

Temperature

What exactly is "hot" and "cold"? Is there a maximum or minimum temperature that an object can reach? Temperature is difficult to define precisely, but we all have an intuitive idea of what is meant by it. It is a measure of a substance's "hotness". Whenever a hot object is placed next to a cold object, the cold one becomes warmer, while the warm one becomes cooler. It is obvious that something is being exchanged here. Is there such a thing as "cold" and "hot" that is being exchanged?

The definition that we will use for *temperature* is "the property of a body or region of space that determines whether or not there will be a net flow of heat into it or out of it from/to a neighboring body or region and in which direction (if any) the heat will flow." Put simply, temperature is the measurement of heat with a thermometer. If there is no heat flow, the bodies or regions are said to be at "thermodynamic equilibrium" and at the same temperature. If there is a flow of heat, the direction of the flow is from the body or regions of higher temperature to the body or region of lower temperature. The scale that is used to measure temperature is arbitrary. The volume of matter varies with temperature and, therefore, can be used to measure temperature. For example, most materials expand when warmed and contract when cooled. Most people are familiar with the mercury thermometer; it is based on the nearly linear expansion and contraction of liquid mercury with changes in temperature.

In the *Celsius* temperature scale (formerly called the "centigrade" scale); the freezing point and the boiling point of water are used as two fixed points for establishing a measurement of temperature. This scale was devised by the Swedish astronomer Anders Celsius (1701 - 1744). He established the freezing point of water as 0 degrees Celsius, and the boiling point of water as 100 degrees Celsius. He then divided the interval between the melting point and the boiling point of water into 100 equal segments. Each of the segments is called one degree Celsius. By extending the same scale divisions beyond the two fixed points, he could measure temperatures below 0 degrees Celsius and above 100 degrees Celsius.

The SI base unit of temperature is the Kelvin (K). In the 19th century, Lord Kelvin proposed a thermodynamic method to specify temperature, based on the measurement of the quantity of heat flowing between bodies at different temperatures. This concept relies on an absolute scale of temperature with an absolute zero of temperature, at which nothing can give up heat. There appears to be no limit as to how hot a substance can get. However, Absolute zero is the lowest temperature that is theoretically attainable. The kinetic energy of atoms and molecules is minimal at the temperature. The unit in which thermodynamic temperature is expressed is the °K. The temperature at Absolute zero is 0 degrees Kelvin ($0^{\circ}K$). The increments of degrees Kelvin are the same as the increments in the Celsius scale. Since

Absolute zero ($0^{\circ}K$) is at a temperature of minus 273.15 degrees Celsius ($-273.15^{\circ}C$), it is easy to convert from the Kelvin scale to the Celsius scale:

$$^{\circ}K = {}^{\circ}C + 273.15 \qquad \text{eq. 1.6}$$

Therefore, a temperature of about 22 $^{\circ}C$, which is about room temperature, would be equal to 295 $^{\circ}K$.

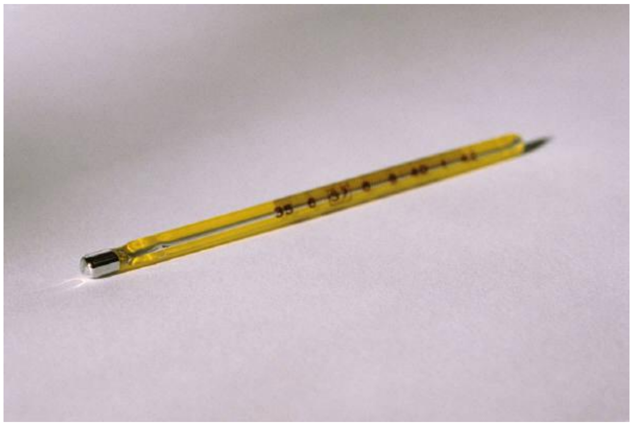
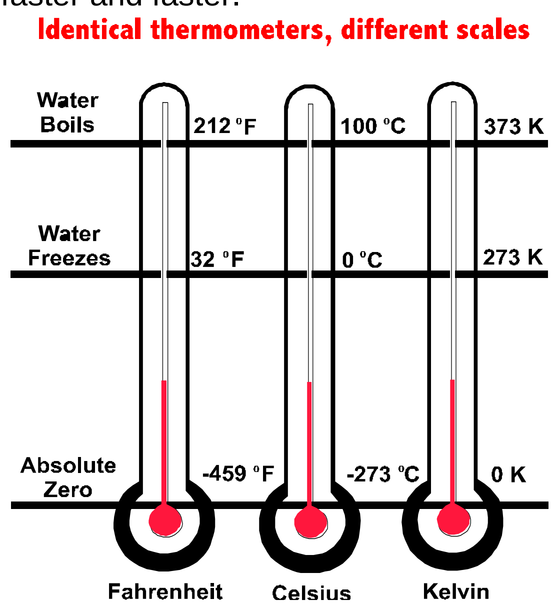


Figure 1.9
The liquid mercury in the bottom well of the glass thermometer expands upon heating. The mercury has no where to go except through the small tube in the thermometer that is marked incrementally, displaying the temperature of the mercury as a relationship to how much it has expanded.

All objects contain internal energy, which consists of potential energy from the interactions between atoms and kinetic energy from random thermal motion of the atoms. Atoms and molecules are constantly vibrating. This vibratory energy is transferred as heat from high to low temperatures. Boiling water will harden an egg as the energy is transferred to the egg, which is trying to reach thermal equilibrium. As the temperature increases, so does the internal, or thermal, energy causing the atoms to move faster and faster.

The Fahrenheit scale is at present the common temperature scale used in the United States. In this scale, the freezing point of water is 32 degrees Fahrenheit ($32^{\circ}F$) and the boiling point of water is 212 degrees Fahrenheit ($212^{\circ}F$). Thus, the number of degrees between the boiling point of water and the freezing point of water is ($212 - 32 = 180^{\circ}F$). Since the number of degrees between the boiling point of water and the freezing point of water in the Celsius scale is 100 degrees, the ratio between Fahrenheit and Celsius is therefore 180 divided by 100 which is equal to 1.8. Therefore, for every degree Celsius there are 1.8 degrees Fahrenheit. The formula to convert degrees Celsius to degrees Fahrenheit is:

$$^{\circ}F = (1.8 \times {}^{\circ}C) + 32 \qquad \text{eq. 1.7}$$



The Common Temperature Scales
Figure 1.10

This formula can be rearranged to get a formula for converting degrees Fahrenheit to degrees Celsius:

$$^{\circ}C = (^{\circ}F - 32) \div 1.8 \qquad \text{eq. 1.8}$$

This equation is sometimes written as $^{\circ}C = (^{\circ}F - 32) \times 5/9$. This equation should be memorized in whichever form is the easiest.

Example: Let us use these formulas to convert 98.6°F, which is a normal human body temperature, to degrees Celsius (°C):

From equation 1.8: $^{\circ}C = (^{\circ}F - 32) \div 1.8$
$^{\circ}C = (98.6 - 32) \div 1.8$
$^{\circ}C = (66.6) \div 1.8$
$^{\circ}C = 37$

Therefore, the body temperature of 98°F is equivalent to 37°C.

Time to think:

7. Convert 25°C to °F______________
____

8. Convert 112 °F to °C. _______________________________.

9. Look up the melting point of any liquid other than water and show that temperature in Celsius and in
Farenheight____.

10. Aluminum melts at 661°C. What is this temperature in degrees F?
____
_________________________________.

11. Aluminum boils at 2520°C. What is this temperature in degrees F?
_________________________________.

You are doing great! Now check your answers.

Number correct: | | Number incorrect: | |

If you made any mistakes go back and make sure you understand the concepts.

Density

Density is defined as the mass of a substance per unit of volume. In SI units, it is measured in kg per cubic meter (kg/m^{-3}). It is called a *derived* physical quantity because it comes because of measuring other physical quantities. The density of any material can be determined by measuring its mass and its volume. The densities of solids are often expressed in grams (g) per cubic centimeter (cc) or Kilograms per cubic meter. The densities of liquids are expressed in grams per cubic centimeter or grams per liter. The densities of gases are expressed in grams per liter.

Density is an important physical property. Knowing the mass of a substance, or knowing the volume of a substance, does not really give us very much information. Many students have been fooled by the question, “Which weighs more a ton of lead or a ton of feathers?” Obviously, a ton is a ton, whether we are speaking of lead, which is very dense, or of feathers. Since the density of the feathers is much less than that of the lead, it will take a greater volume of feathers to equal the same mass as the lead. Students should prepare for the other trick question that usually follows the above question: "Which weighs more, an ounce of lead or an ounce of gold?" Be aware that gold is measured in troy ounces and lead is measured in avoirdupois (pronounced “av ’er de poiz’ “) ounces (so the answer is “gold”).

Finding the masses of equal volumes of different substances will give us the density of the substances. The densities of specific substances are especially useful when comparing them to the densities of other substances. Density tables for many substances can be found in many reference books including a “CRC Handbook” or, for elements, some periodic tables. At one time, the term "specific gravity" was used to show the relative density of a substance as compared to the density of water. The density of water at 4^oC is one gram per cubic centimeter (1.0 g/cc). A cc of water just happens to be the same volume as one ml of water. Therefore, the density of water at 4^oC is 1g/ml.

Peaceful
“When you find peace within yourself, you become the kind of person who can live at peace with others”.
-Peace Pilgrim

☺
Q: What happens when electrons lose their energy?
A: They get Bohr'ed.
☺

A Laboratory Experiment: the Density of Water and Ice Lab 1.2

In this experiment, we calculate the exact density of ice and water.

The purpose of this experiment is to demonstrate that the density of ice (a solid) is less than that of the liquid water. To discover how the density of a substance is an important identification trait. The density of ice and water are calculated.

The apparatus required:

- One small beaker
- Hot plate
- Five ml graduated pipette
- 10 ml graduated cylinder
- Triple beam balance
- Freezer

Procedure:

1. Allow some water to come to a boil in a beaker on a hot plate in order to aid in removing air bubbles.
2. Allow the water to cool to room temperature.
3. Find the weight of the empty 5ml pipette.
4. Record the temperature of the water and add about 4 ml (record the volume to +/- 0.1 ml) to the 5 ml pipette.
5. Find the weight of the pipette and water.
6. Put the pipette with water into the freezer overnight.
7. The next day, find the new volume of the ice in the pipette.
8. Find the weight of the ice and pipette.
9. Allow the ice to melt by warming the pipette with the hands.
10. Record the volume and weight and compare to the previous day.
11. Use the chart (table 1.4) below to fill in the data. (Remember that density is calculated by dividing the mass of the substance by the volume of the substance).

Density of water and ice Laboratory	**Table 1.4**
Mass of empty pipette	
Mass of water and pipette	
Mass of water alone	
Volume of water	
Volume of ice	
Mass of ice and pipette	
Mass of ice alone	
Density of ice	
Density of water	

12. Write the conclusions in your own words.

____
____
____
____
____
____
____
____
____
____
____
____
____

The Relative Density Demonstration *Demo 1.1*

Only the instructor should conduct this demonstration.

The purpose of this demonstration is to show how different substances in the same physical states may have very different densities.
Supplies required:

- One 500 ml graduated cylinder or other appropriate container
- Approximately 100 ml of liquid mercury
- A couple of metal nuts or bolts small enough to fit into the graduated cylinder
- Approximately 100 ml of carbon tetrachloride or another liquid of the same density

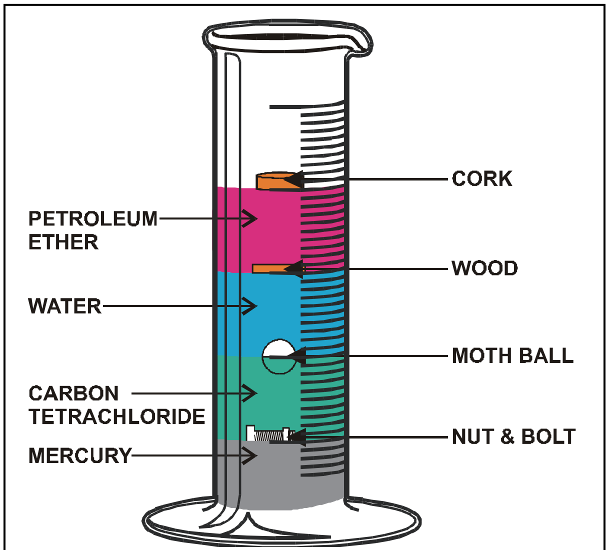


Figure 1.11 Relative Densities Demonstration

- One or two mothballs
- 100 ml of water
- Food coloring, if available
- A small piece of wood with a density just below that of water
- Approximately 100 ml of petroleum ether or an oil of equal density
- A small cork

Procedure:
1. To a 500 ml graduated cylinder, add 100 ml of mercury.
2. Add a couple of the metal nuts or bolts onto the surface.
3. Add 100 ml of carbon tetrachloride and add one or more mothballs to the surface.
4. Add 100 ml of water (food coloring added to the water is a good effect) and place a piece of wood on the surface. (Try adding a few nail brads or tacks to get the proper density for the wood piece).
5. Add 100 ml of petroleum ether and put a cork on the surface.

Note: If other substances besides carbon tetrachloride or petroleum ether are used, be sure that they are immiscible in water. In addition, all of these liquids (except for water) are potentially dangerous and should be handled carefully and inside a fume hood or with other adequate ventilation. The wood may need to be weighted down with tacks or nails in order to keep it floating above the water and below the oil.

Time to think:

12. Why is the piece of wood suspended between the layers of petroleum ether and water?
____

13. Where can the density of mercury be found?
____

14. How could the density of the metal bolt be determined?

15. The substances with densities between those of carbon tetrachloride and cork are:

You are on your way towards completing Unit One!

Now check your answers.
If you made any mistakes go back and make sure you understand the concepts.

Chapter Two Review

Teacher's Initials ________________

1. What does the word "philosopher" mean?

2. In your personal opinion, why do you think the atomists are worth remembering? (Teacher initial answer)

3. What is the difference between an atom and an element?

4. Discuss something about the scientist John Dalton that you find interesting. (Teacher initial answer)

5. How does the Thomson model of the atom differ from Bohr's model (Bohr's Theory) of the atom?

6. If a diamond is pure carbon, and graphite is pure carbon, why do you think their properties are so different?

7. What is the difference between heat and temperature?

8. Can you think of a method to measure something at absolute zero? (The presence of a measuring device, such as a thermometer, would heat the substance).

9. In the demonstration of relative densities, how could it be that a metal bolt can float?

10. The last Fourth of July many people at the California beaches enjoyed ninety-degree (Fahrenheit) weather. Express this temperature in degrees Celsius. Convert this temperature to degrees Kelvin.

11. How many cubic centimeters (cc) are in a cubic yard (1 yd^3)?

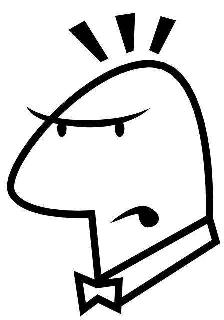
12. How much volume would 2 pounds of water occupy?

Now check your answers.

Number correct: Number incorrect:

Which drawing below represents peaceful?





Chapter Three

Uncertainty

Can one ever be totally certain of anything? Measurements always involve a comparison and some uncertainty. When a particular book measures 30 cm long, we are actually saying that the book is 30 times as long as an object that is one cm long. Since we normally do not carry objects with us that are exactly one centimeter long (or 1 inch long), we have invented the yardstick and the meter stick to aid us in our measurements. In ancient times, measurements were made according to objects that were actually carried on the body, ready to use. For instance, the cubit of biblical time was the length of a man's forearm or the distance from the tip of the elbow to the end of his middle finger (about 60 cm). The cubit as a unit of measuring length was used in the account of Noah and the Ark: "Genesis 6:14, 15 Make thee an ark of gopher wood; rooms shalt thou make in the ark, and shalt pitch it within and without with pitch. And this is the fashion which thou shalt make it of: The length of the ark shall be three hundred **cubits**, the breadth of it fifty **cubits**, and the height of it thirty **cubits"**. This was useful, because a man's forearm was readily available, convenient, and could not be mislaid. However, it was not a positive fixed dimension or a standard. Even though the cubit is no longer used as a unit of measurement, other customary standards originated in the same way. Our foot-rule started out as the length of a man's foot. Consequently, in the early days of history the foot varied as much as 3 or 4 inches in length. The Egyptians and the Greeks used a wheat seed as the smallest unit of weight. This standard was uniform and accurate for the times. The grain is still in limited use as a standard weight (look at the label on most bottles of aspirin).

Fig. 1.12 The unit of "cubit" for Noah's ark

It is inaccurate to say that anything measures *exactly* something unless we are defining the standard. One could say a book is exactly "one book" long. At this time, I could establish my own measurement system to where everything else is measured according to the length of that particular book.

Comparison always involves some uncertainty. *Absolute uncertainty* is an expression of the margin of uncertainty associated with the measurement. If the estimated uncertainty in reading a perfectly calibrated ruler is plus or minus (+/-) 0.1 cm, then we call the quantity "plus or minus 0.1 cm", the *Absolute uncertainty* associated with the reading. Absolute uncertainty always has the same units as the measurement.

Relative uncertainty is an expression comparing the size of the absolute uncertainty to the size of its associated measurement. It is dimensionless:

$$\text{Relative uncertainty} = \frac{\text{Absolute uncertainty}}{\text{Magnitude of measurement}} \qquad \text{eq. 1.9}$$

The *percent relative uncertainty* is obtained by multiplying the relative uncertainty by 100:

Percent relative uncertainty = Relative uncertainty X 100

Therefore, in the above example of measuring the book, the relative uncertainty is: +/- 0.1 cm /30cm = 0.0033. And the percent relative uncertainty = 0.0033 X 100 = 0.33%

Propagation of Error

Whenever two or more quantities are measured directly in order indirectly to determine the value of another, the question of propagation of errors must be addressed. Consider the example of the measuring of the book. If we have three books lying end to end on the table, and measure each book individually, in order to determine the total length of the books by adding the three measurements, we may have a problem of *propagation of error*. If our meter stick can measure each book to a precision of +/-1.0 mm, then each measurement could be up to 1.0 mm off. Now, we could say that the maximum error would be 3.0 mm. However, this is not very accurate because each measurement could be 1.0 mm larger or 1.0 mm smaller than the actual measurement. It is possible that all the errors in measurement could cancel each other out. Luckily, mathematicians have come to our rescue. The propagated uncertainty is handled this way; each absolute uncertainty is given a designation e_1, e_2, e_3, …. . In our above example, each of these is 0.1 cm (1.0 mm). When adding or subtracting uncertainties, the total absolute uncertainty (e_t) is calculated as:

From the example; $e_t = (e_1^2 + e_2^2 + e_3^2)^{1/2}$ eq 1.10

[$(x)^{1/2}$, means that we are taking the square root of x]

$e_t = (1.0^2 + 1.0^2 + 1.0^2)^{1/2}$

$e_t = (3)^{1/2} =$ 1.73

Thus, the total absolute uncertainty for this propagated uncertainty is 1.73 mm.

When the uncertainties require multiplication or division, we must find the percent relative uncertainty. All uncertainties must first be converted to percent relative uncertainties. Then calculate the error of the product or quotient as follows:

$\%e_t = (\%e_1^2 + \%e_2^2 + \%e_3^2)^{1/2}$ eq. 1.11

Time to think:

1. Why do you think the cubit is not used as a system of measure in modern times?

____

____

2. If the absolute uncertainty in a measurement is 1 ml and the measurement is 902 ml, what is the relative uncertainty?

____

3. What is the percent relative uncertainty for the above figures?

____

4. Use equation 1.10 to calculate the total absolute uncertainty for three measurements each with an uncertainty of 2.0mm.

____

____

You are doing great! Now check your answers.

Number correct: ☐ Number incorrect: ☐

If you made any mistakes go back and make sure you understand the concepts.

Mean and Standard Deviation

When we encounter mixed operations, such as addition and/or subtraction in the same equation as multiplication and or division, we must first work out the addition and subtraction portion of the problem in absolute uncertainties and then convert to relative uncertainties.

Whenever a scientist gathers several bits of data, he or she will find it convenient to deal with only one number. It is usually best to calculate the arithmetic average of the numbers. This average is called the *mean*. The *standard deviation* measures how closely the data (the numbers) are clustered about the mean. The smaller the standard deviation, the more closely the data are clustered about the mean. The standard deviation is rather like the "mean of the mean," and often can help one find the story behind the data. Each deviation from the mean is calculated by subtracting the measured value from the mean. An example from a biochemistry lab may aid in seeing the importance of these concepts.

Suppose we are measuring the effect of temperature on the rate of reaction of an enzyme. Using a stopwatch, we get four different measurements: 82.1 seconds, 78.3 seconds, 85.5 seconds, and 83.4 seconds. To find the average (x) and standard deviation (s):

x = (82.1s + 78.3s + 85.5s + 83.4s) ÷ 4

x = 82.3 s

$$s = \sqrt{\frac{(e1 - x)^2 + (e2 - x)^2 + (e3 - x)^2 + (e4 - x)^2}{(n - 1)}}$$ General formula eq. 1.12

n = number of measurements

Equation 1.12 becomes:

$$s = \sqrt{\frac{(82.1 - 82.3)^2 + (78.3 - 82.3)^2 + (85.5 - 82.3)^2 + (83.4 - 82.3)^2}{(4 - 1)}}$$

s = 3.03 s

The significance of the standard deviation is that it measures the width of the error curve, as shown in figure 1.13 and figure 1.14.

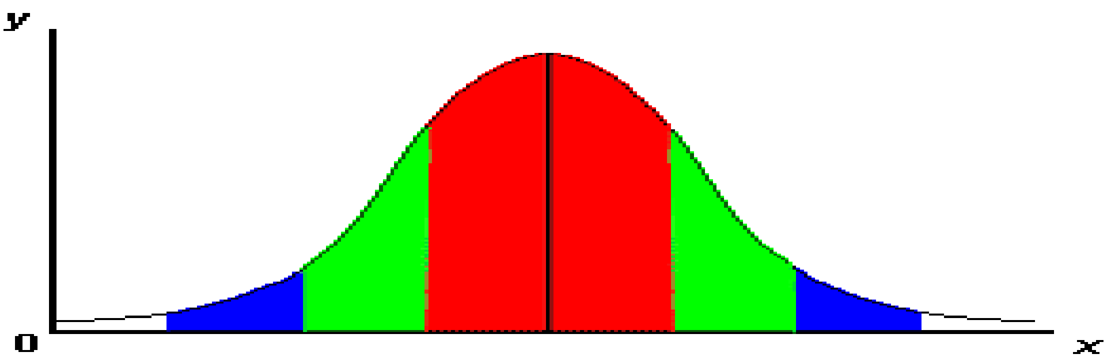


Figure 1.13

The wider the bell curve, the larger the standard deviation (fig. 1.13). With a narrow bell curve, the standard deviation will be smaller (fig. 1.14).

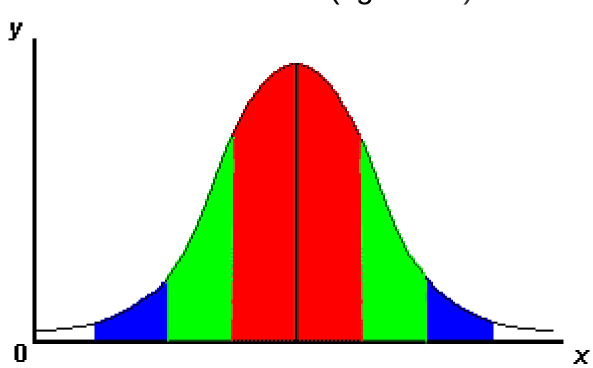


Figure 1.14

Logarithms

Two kinds of *logarithms* are often used in chemistry: common (or Briggian) logarithms and natural (or Napierian) logarithms. The power to which a base of 10 must be raised to obtain a number is called the common logarithm (log) of the number. The power to which the base e (e = 2.718281828.......) must be raised to obtain a number is called the natural logarithm (ln) of the number. Earlier we introduced scientific notation. To put it simply, logs are exponents. In operations with logarithms, we follow the law of exponents. A logarithmic scale is a scale of measurement in which an increase or decrease of one unit represents a tenfold increase or decrease in the quantity measured. The Richter scale, decibels, and pH measurements are common examples of logarithmic scales of measurement. The use of logarithms allows much time saving in multiplication, division, and other operations. Some of the properties of logarithms are:

$b^x \times b^y = b^{x+y}$ eq.1.13
$b^x / b^y = b^{x-y}$ eq. 1.14
$(b^x)^n = b^{xn}$ eq. 1.15
$\sqrt[n]{b^x} = b^{x/n}$ eq.1.16

Using a previous example in this unit, we saw that 100 could be written as 10 x 10 or 10^2. Using logarithms, $\log_{10}$ ("log to the base 10"); $\log_{10}100 = 2$ is equivalent to $10^2 = 100$. Alternatively, in plain English, the log (in base ten) of one hundred equals two. The log of any real number can be calculated. The easiest way is to just put the number into a calculator and push the “log” key. The number does not have to be an even power of ten. For instance, the log of 67 is 1.826. This should make sense. The log of 10 is 1, and the log of 100 is 2. Therefore, the log of 67 should be somewhere between one and two, since 67 is somewhere between ten and one hundred.

The natural logs are little tougher to understand. Many equations used in chemistry were derived using calculus, and these often-involved natural logarithms. The relationship between ln x and log x is:

ln x = 2.303 log x. eq. 1.17

Many important formulas in calculus take their simplest possible forms using the natural logarithms.

Practice:

Using logarithms, calculate the speed of an airplane, in meters per seconds, that has traveled 100,000 meters in 1000 seconds. Although this problem is easy to do in our head, let us try it using logarithms. In scientific notation, 100,000 equals 1 x 10^5, and 1000 equals 1 x 10^3. Using logarithms, the log to the base 10 of 100,000 is 5, the log to base 10 of 1000 is 3. Using the law of logarithms for division, equation 1.14, the answer is 5 minus 3 equals 2. The *antilog* of 2 is 100, so the answer is 1 x10^2 m/s or **100 m/s**.

Time to think:

Use the rules of logarithms to solve the following:

(There is no need to use a calculator)

5. $10^2 \times 10^4$ = ______________________
6. 5^3 X 5^6 = ________________________
7. A^4 X A^2 = _________________________
8. 10,000,000/100 = __________________
9. $(6^4)^2$ = ___________________________
10. $(10^3)^5$ = __________________________
11. $\sqrt[2]{3^4}$ = ___________________________
12. Y^{10}/Y^4 = _________________________

You are doing great! Now check your answers.

Number correct: Number incorrect:

If you made any mistakes go back and make sure you understand the concepts.

Avagadro's Constant

Avagadro's constant, formerly known as Avagadro's number, is defined as the number of atoms or molecules in one *mole* of the substance. Basically, it is just a very big number that enables chemists to use practical amounts of substances. How many doughnuts are in a dozen doughnuts? How many eggs are in a gross of eggs? Sometimes it is easier to talk about one egg, such as when you are making yourself breakfast, sometimes it is easier to talk about a gross, such as when buying eggs for your restaurant. In the laboratory, chemists do not deal with one atom or molecule. Chemists normally use quantities of substances that can be seen and measured. At the same time, the chemist in the laboratory is often concerned with exact amounts of substances involved in the reactions. Thus, the chemist is concerned with how many atoms or molecules are reacting, but on a relatively large scale.

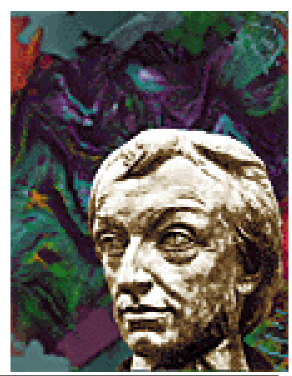


Fig. 1.15 Amedeo Avagadro (1776- 1856)

Amedeo Avagadro is the Italian chemist and physicist responsible for Avagadro's constant. He did not actually invent the number that is named after him. Avagadro derived the cocstant from *Avagadro's law*, which he stated in 1811. Avagadro's law states that equal volumes of gases all contain equal numbers of molecules at the same pressure and temperature. This “law” is true only for ideal gases. This law, (which was actually a hypothesis) provided a method of calculating molecular weights from vapor densities. Professor Avagadro produced a four-volume work of physics between 1837 and 1841. Avagadro received no recognition in his lifetime. The prime reason for this was his lack of experimentation to verify his hypotheses. He was vindicated when a colleague presented a system of atomic weights that was determined by using Avagadro's work. Amedeo Avagadro is now considered one of the most important chemists and physicists of the 19th-century.

Avagadro's constant was determined by calculating the density of 22.4 liters (one mole) of a gas at zero degrees Celsius and one atmosphere of pressure. After calculating the density, estimating the size of the atoms, and knowing the volume, early scientists were able to estimate the number of molecules, or atoms in a mole. The gas in 22.4 liters was found to contain 602,000,000,000,000,000,000,000 atoms. This is equal to 602 billion trillion atoms. This number is one *mole* of atoms or molecules. The number has been modified to be defined as the number of carbon (C_{12}) atoms in exactly 12.00 g of carbon. This number is a constant and is equal to 6.022×10^{23}. Therefore, a mole is just a number, like a dozen or a gross. A mole of anything is equal to 6.022×10^{23}. A mole (abbreviated mol) of any ideal gas will occupy 22.4 liters of space at standard temperature and pressure. A mole of hydrogen molecules is equal to 6.022×10^{23} molecules. A mole of methane (CH_4) molecules is equal to 6.022×10^{23} molecules. A mole of eggs is equal to 6.022×10^{23} eggs. Therefore, the number of atoms or molecules in one mole of the substance is equal to Avagadro's number, which is 6.022×10^{23}. This is a very large number. It is hard for us to imagine just how large this number is. If someone gave you exactly 12 g of carbon and asked you to count the number of atoms in that amount, and you can count them at a rate of one atom per second, it would take you approximately 10^{19} years to count them all. This is about ten thousand times the age of the universe. If each of these atoms in one mole of carbon were the size of a softball, the resulting volume would be about the one-half the size of the earth. However, atoms are quite small and a mole of carbon atoms would approximately fit into a tablespoon. This is why chemists use the mole; it is a manageable quantity of atoms or molecules.

Figure 1.16

The mass of one mole of any substance will vary according to the mass of one unit of the substance. The atomic weights of the elements are listed in the periodic table. Figure 1.16 shows the element of carbon as it appears in the periodic table. The number below the symbol for carbon (C) is 12.0107. This number represents the mass in grams of one mole of naturally occurring carbon atoms. The reason that the number is not exactly 12.00 g/mol is because in nature, some of the carbon atoms contain an extra neutron. This added mass is averaged out. Since most of the carbon atoms weigh exactly 12 g per mole, and some of the carbon atoms have a mass of 13 g per mole (the added mass due to the extra neutron), the result is such that 6.022×10^{23} atoms have a total mass of 12.0107 grams. In figure 1.16, the figure that appears above the atomic symbol for carbon is the atomic number. In this case, the number is six. This number represents the number of protons in the nucleus. The number of protons is what defines an element. A carbon twelve (C_{12}) atom has six protons and six neutrons. In nature, some of the carbon atoms have seven neutrons. The carbon thirteen (C13) atoms have an atomic weight of 13 g per mole. The periodic table will be fully explored in Unit Two.

The Periodic Table of the Elements

Period	1 IA 1A	2 IIA 2A	3 IIIB 3B	4 IVB 4B	5 VB 5B	6 VIB 6B	7 VIIB 7B	8	9	10	11 IB 1B	12 IIB 2B	13 IIIA 3A	14 IVA 4A	15 VA 5A	16 VIA 6A	17 VIIA 7A	18 vIIIA 8A
1	1 H 1.008																	2 He 4.003
2	3 Li 6.941	4 Be 9.012											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
3	11 Na 22.99	12 Mg 24.31											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.07	17 Cl 35.45	18 Ar 39.95
4	19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.88	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.47	28 Ni 58.69	29 Cu 63.55	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.92	34 Se 78.96	35 Br 79.90	36 Kr 83.80
5	37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.94	43 Tc (98)	44 Ru 101.1	45 Rh 102.9	46 Pd 106.4	47 Ag 107.9	48 Cd 112.4	49 In 114.8	50 Sn 118.7	51 Sb 121.8	52 Te 127.6	53 I 126.9	54 Xe 131.3
6	55 Cs 132.9	56 Ba 137.3	57 La* 138.9	72 Hf 178.5	73 Ta 180.9	74 W 183.9	75 Re 186.2	76 Os 190.2	77 Ir 190.2	78 Pt 195.1	79 Au 197.0	80 Hg 200.5	81 Tl 204.4	82 Pb 207.2	83 Bi 209.0	84 Po (210)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra (226)	89 Ac~ (227)	104 Rf (257)	105 Db (260)	106 Sg (263)	107 Bh (262)	108 Hs (265)	109 Mt (266)	110 Uun ()	111 Uuu ()	112 Uub ()		114 --- ()		116 --- ()		118 --- ()

58 Ce 140.1	59 Pr 140.9	60 Nd 144.2	61 Pm (147)	62 Sm 150.4	63 Eu 152.0	64 Gd 157.3	65 Tb 158.9	66 Dy 162.5	67 Ho 164.9	68 Er 167.3	69 Tm 168.9	70 Yb 173.0	71 Lu 175.0
90 Th 232.0	91 Pa (231)	92 U (238)	93 Np (237)	94 Pu (242)	95 Am (243)	96 Cm (247)	97 Bk (247)	98 Cf (249)	99 Es (254)	100 Fm (253)	101 Md (256)	102 No (254)	103 Lr (257)

Figure 1.17

The important thing to remember is that in chemistry, we count by the masses of the substances. Looking at the periodic table, we find that one mole of carbon weighs 12.01 g per mole. We also see that one mole of hydrogen weighs 1.008 g per mole. Thus, we can calculate the *molecular weight* of one mole of methane (CH_4). One mole of methane consists of one mole of carbon and four moles of hydrogen. Thus, the molecular weight of methane is:

1 x 12.01 g/mol	carbon
+ 4 x 1.008 g/mol	hydrogen
16.042 g/mol	methane

It is important to see the ratios in this equation. In any amount of methane, the ratio of the mass of carbon to the ratio of the mass of hydrogen is approximately 12 to 4 (which is the same as 3:1). If we were forming methane from carbon, we had 6 g of carbon and excess hydrogen, and all of the carbon reacted to form methane, then two grams of hydrogen will have reacted to form methane. This concept is very important because knowing the ratios in which substances react, is a very important aid in determining the reactants and the products.

The *molecular weight*, or more appropriately the *relative molecular mass*, is defined as "the ratio of the average masses per molecule of the naturally occurring form of an element or compound to 1/12 the mass of carbon twelve (C_{12}) atoms". It is equal to the sum of the relative atomic masses of all the atoms that comprise a molecule. The molecular weight is similar to the atomic weight, except that the molecular weight applies to the entire molecule. The molecular weight of any particular molecule is the sum of the atomic weights of all the atoms in the molecule.

Time to Think:

13. How many molecules are in one mole (mol) of water (H_2O)?

____

14. How does inorganic chemistry differ from organic chemistry?

____

15. How does molecular weight differ from atomic weight?

____

You are doing great! Now check your answers.

Number correct: __________ Number incorrect: ___________

If you made any mistakes go back and make sure you understand the concepts.

Chapter Three Review

Teacher's Initials ________

1. If you were to start your own system of measurement of lengths, weights, and volumes, how would you begin? On what would you base your measurements? (Teacher initial)

____

____

2. What is the difference between relative uncertainty and absolute uncertainty?

____

3. Which branch of chemistry seems the most interesting to you? Why? (Teacher initial)

____

____

4. If you were investigating the plating techniques for printed circuit boards, which branch of chemistry would you be studying?

____

5. How are biology and biochemistry related?

____

6. Why wasn't Amedeo Avagadro recognized as a profound chemist in his lifetime?

____

7. Which quantity contains the most marbles: a dozen, a gross, a mole, or a score of marbles?

____

8. Given the following atomic weights, calculate the molecular weight of water (H_2O).

   H = 1.008 g/mol; O = 16.00 g/mol.

____

____

You are doing great! Now check your answers.

Number correct: ☐ Number incorrect: ☐

Unit One Review

Date: ___________ Score:_________________

Four points each

May use a calculator, periodic table, table 1.3 (below)
and equation 1.4; $Q = s \times m \times \Delta T$

SAE	Metric
1 inch	2.54 cm
1 mile	1.609 km
1 gallon	3.75×10^{-3} m^3
1 atmosphere	1.013×10^5 Pascal
1 KWh	3.6×10^6 Joule
1 ounce	28.35 grams
1 quart (liq.)	0.94635 liter

Table 1.3 Selected conversion units

1. How does a scientist disprove a hypothesis?

____

____

2. Express the following in scientific notation: (one point each)

    a. 10,100____

    b. 0.010100____

    c. 2,000.00300____

    d. One thousandth____

3. How many mm are in a yard? ______________________________________

4. The Captain of an ocean liner tells the passengers that the iceberg being observed in the distance is only the “tip” of the iceberg. He states that the iceberg is still 2.5 km away. How many feet away is the iceberg?

____

5. How many thermal calories (c) are in 100 food Calories (C)?

____

6. State which is larger

a. an atom of oxygen

b. a molecule of oxygen

7. Rank the following from coldest to warmest temperature;

212° F, 80° C, 180° K. _______________________________________

8. A large, heavy, watertight box has been delivered to your home. What measurements must be made to determine if it is denser than water without actually putting it in water?

____

____

9. Express room temperature (25° C) in degrees Fahrenheit____________

____

____

10. Give an example of an organic compound.

____

11. Find the average of the following numbers: 220.4, 222.6, 225.0, 223.8

____

12. Which will contain more atoms: a mole of copper or a mole of gold?

____

13. Which will have more mass, a mole of potassium or a mole of phosphorus?

____

14. How many watts are in ten gigawatts?

____

15. Given the following atomic weights, calculate the molecular weight of sodium hydroxide (NaOH); O = 16.00 g/mol, Na = 22.99 g/mol, H = 1.008 g/mol.

____

16. Which has more kinetic energy?

a. a solid

b. a liquid

c. a gas

17. Express 10,000 grams in terms of pounds.

____
____
____

18. The specific heat capacity of silver is .235 J/g C°. How much energy is required to heat 30 grams of silver from 25°C to 125°C?

____
____

19. How much heat is given off when 5.0 grams of silver cools from 100°C to 40°C? (Specific heat capacity is found in problem # 18). __________________

____
____

20. What is the volume (in ml) of 2 pounds of a liquid with a density of 1.0 grams/ml?

____
____

21. Which has more atoms?

a. A mole of hydrogen

b. A mole of sodium

c. A mole of H_2O

d. A mole of NaCl

e. none of the above

22. Equation 1.3 states KE = ½ mv^2. If the system in question is a ball rolling down a hill, which will increase the kinetic energy the most?

a. double the mass of the ball

b. double the velocity of the ball

c. double the temperature

d. none of the above, the system is in equilibrium

23. How many mm^3 are in one cubic meter?

____

____

24. Silver melts at 960.8 °C and has a specific heat capacity of .235 J/g C°. What is the melting point of silver in °F?

____

____

25. The density of mercury is 13.6 g/ml. How many milliliters are there in 500 g of mercury?

____

____

Number correct: ☐ Number incorrect: ☐

Congratulations! You have completed Unit One

CHEMISTRY UNIT ONE TEST

Name: ________________________________ **Date:** ________ **Score:** _______

May use a calculator, periodic table, table 1.3 (below)
and equation 1.4; Q = s x m x ΔT

SAE	Metric
1 inch	2.54 cm
1 mile	1.609 km
1 gallon	3.75×10^{-3} m^3
1 atmosphere	1.013×10^5 Pascal
1 KWh	3.6×10^6 Joule
1 ounce	28.35 grams
1 quart (liq.)	0.94635 liter

Table 1.3 Selected conversion units

1. How can a theory be proven?

____

____

2. Express the following in scientific notation: (one point each)

    a. 100,100____

    b. 0.0010100____

    c. 200.00300____

    d. One hundredth____

3. How many mm are in a foot? ____________________________________

4. The Captain of an ocean liner tells the passengers that the iceberg being observed in the distance is only the "tip" of the iceberg. He claims that 91.7% of the iceberg is under water. Using this information and assuming the ocean to be pure water (of

course in reality it has salt in it) with a density of 1.000 grams per ml, find the density of the ice in g/ml to 3 decimal places.

____

____

5. How many thermal calories (c) are in 1000 food Calories (C)?

____

6. State which is larger

a. an atom of oxygen

b. a molecule of oxygen

7. Rank the following from coldest to warmest temperature;

32° F, 20° C, 200° K. ____

8. A large, heavy, water-tight box has been delivered to your home. What measurements must be made to determine if it will float in your pond without actually putting it in water?

____

____

9. Express the outside temperature (assume 27° C) in degrees

Fahrenheit____

____

10. Give an example of an organic compound.

____

11. Find the average of the following numbers: 322.8, 330.1, 331.4, 328.2

____

12. Which will contain more atoms; a mole of copper or a mole of gold?

____

13. Which will have more mass, a mole of silicon or a mole of sulfur?

____

14. How many watts are in a gigawatt?

____

15. Given the following atomic weights, calculate the molecular weight of ammonia (NH_3); N = 14.01 g/mol, H = 1.008 g/mol.

____

16. Which has more kinetic energy?

a. a solid

b. a liquid

c. a gas

17. Express 5,000 grams in terms of pounds.

____
____
____

18. The specific heat capacity of silver is .235 J/g C°. How much energy is required to heat 30 grams of silver from 25°C to 75°C?

____
____

19. How much heat is given off when 100 grams of silver cools from 100°C to 40°C? (Specific heat capacity is found in problem # 18). ___________________

____
____

20. What is the volume (in ml) of 1 pound of a liquid with a density of 0.9 grams/ml?

____
____

21. Which has more atoms?

a. A mole of hydrogen

b. A mole of oxygen

c. A mole of H_2O

d. A mole of NaCl

e. none of the above

22. Equation 1.3 states KE = ½ mv^2. If the system in question is a ball rolling down a hill, which will increase the kinetic energy the most?

a. Triple the mass of the ball

b. Double the velocity of the ball

c. Quadruple the temperature

d. none of the above, the system is in equilibrium

23. How many cm^3 are in one cubic meter?

____

____

24. Silver melts at 960.8 °C and has a specific heat capacity of .235 J/g C°. What is the melting point of silver in °F?

____

____

25. The density of mercury is 13.6 g/ml. How many milliliters are there in 1.0 kg of mercury?

____

____