The Chemistry Resource Book

Information for Elementary and Middle School Teachers

Third Edition

2009

Barbara A. Gage, Editor

hydrogen 1 H 1.0079																		helium 2 He 4.0026
lithium 3 Li 6.941	beryllium 4 Be 9.0122			Key: element name atomic number symbol atomic weight (mean relative mass)									boron 5 B 10.811	carbon 6 C 12.011	nitrogen 7 N 14.007	oxygen 8 O 15.999	fluorine 9 F 18.998	neon 10 Ne 20.180
sodium 11 Na 22.990	magnesium 12 Mg 24.305												aluminium 13 Al 26.982	silicon 14 Si 28.086	phosphorus 15 P 30.974	sulfur 16 S 32.065	chlorine 17 Cl 35.453	argon 18 Ar 39.948
potassium 19 K 39.098	calcium 20 Ca 40.078		scandium 21 Sc 44.956	titanium 22 Ti 47.867	vanadium 23 V 50.942	chromium 24 Cr 51.996	manganese 25 Mn 54.938	iron 26 Fe 55.845	cobalt 27 Co 58.933	nickel 28 Ni 58.693	copper 29 Cu 63.546	zinc 30 Zn 65.39	gallium 31 Ga 69.723	germanium 32 Ge 72.61	arsenic 33 As 74.922	selenium 34 Se 78.96	bromine 35 Br 79.904	krypton 36 Kr 83.80
rubidium 37 Rb 85.468	strontium 38 Sr 87.62		yttrium 39 Y 88.906	zirconium 40 Zr 91.224	niobium 41 Nb 92.906	molybdenum 42 Mo 95.94	technetium 43 Tc [98]	ruthenium 44 Ru 101.07	rhodium 45 Rh 102.91	palladium 46 Pd 106.42	silver 47 Ag 107.87	cadmium 48 Cd 112.41	indium 49 In 114.82	tin 50 Sn 118.71	antimony 51 Sb 121.76	tellurium 52 Te 127.60	iodine 53 I 126.90	xenon 54 Xe 131.29
caesium 55 Cs 132.91	barium 56 Ba 137.33	57-70 *	lutetium 71 Lu 174.97	hafnium 72 Hf 178.49	tantalum 73 Ta 180.95	tungsten 74 W 183.84	rhenium 75 Re 186.21	osmium 76 Os 190.23	iridium 77 Ir 192.22	platinum 78 Pt 195.08	gold 79 Au 196.97	mercury 80 Hg 200.59	thallium 81 Tl 204.38	lead 82 Pb 207.2	bismuth 83 Bi 208.98	polonium 84 Po [209]	astatine 85 At [210]	radon 86 Rn [222]
francium 87 Fr [223]	radium 88 Ra [226]	89-102 **	lawrencium 103 Lr [262]	rutherfordium 104 Rf [261]	dubnium 105 Db [262]	seaborgium 106 Sg [266]	bohrium 107 Bh [264]	hassium 108 Hs [269]	meitnerium 109 Mt [268]	ununnilium 110 Uun [271]	unununium 111 Uuu [272]	ununbium 112 Uub [277]		ununquadium 114 Uuq [289]				

*lanthanoids	lanthanum 57 La 138.91	cerium 58 Ce 140.12	praseodymium 59 Pr 140.91	neodymium 60 Nd 144.24	promethium 61 Pm [145]	samarium 62 Sm 150.36	europium 63 Eu 151.96	gadolinium 64 Gd 157.25	terbium 65 Tb 158.93	dysprosium 66 Dy 162.50	holmium 67 Ho 164.93	erbium 68 Er 167.26	thulium 69 Tm 168.93	ytterbium 70 Yb 173.04
**actinoids	actinium 89 Ac [227]	thorium 90 Th 232.04	protactinium 91 Pa 231.04	uranium 92 U 238.03	neptunium 93 Np [237]	plutonium 94 Pu [244]	americium 95 Am [243]	curium 96 Cm [247]	berkelium 97 Bk [247]	californium 98 Cf [251]	einsteinium 99 Es [252]	fermium 100 Fm [257]	mendelevium 101 Md [258]	nobelium 102 No [259]

STEM Resource Center
Prince George's Community College
2009

PRINCE GEORGE'S
COMMUNITY COLLEGE

PRINCE GEORGE'S COUNTY PUBLIC SCHOOLS
CHILDREN COME FIRST

NSF

The Chemistry Resource Book

Information for Elementary and Middle School Teachers

Third Edition

Barbara A. Gage, Editor

June 2009

This book is a publication of the STEM Resource Center of Prince George's Community College. Its development was supported in large part by a grant from the Maryland Higher Education Commission, Dwight D. Eisenhower Mathematics and Science Education Act, Grant Number 94-556. This substantially revised edition was funded by NSF Math Science Partnership Grant #0831970.

TABLE OF CONTENTS

PREFACE

The Chemistry Resource Book in intended for use in science workshops for elementary and middle school teachers offered through the STEM Resource Center of Prince George's Community College. It may also serve as a reference for content and useful laboratory experiences for anyone teaching chemistry at the pre-high school level.

The content of this book is organized into chapters. Each chapter deals with a different chemistry concept. These concepts have been selected because they are the "big ideas" in chemistry and provide a basis for further learning in the science. Furthermore, it is anticipated that a grasp of these concepts provides a foundation for teaching the curriculum of most elementary and middle school chemistry topics. While each chapter may stand alone, it is suggested that the concepts be studied sequentially as the concepts build one upon the other.

Since teaching science involves more than just knowing the material, and it has been shown that students learn science better if they have an opportunity to 'discover' it for themselves, each topic begins with "guided inquiry" hands-on activities. These laboratory activities can be employed as whole-class demonstrations, small group, or individual laboratory experiences. An effort has been made to include activities that require materials and equipment available to most science classrooms, and in some cases, alternatives are suggested. Adherence to laboratory safety regulations is an uppermost consideration. These activities are followed by an elaboration of the concepts. This sequence of activities followed by discussion and practice is in keeping with current research findings about how to best structure instruction so that the construction of learners' knowledge is facilitated. Also included are suggestions for warm-up exercises, thought-stimulating questions, and possible student 'misconceptions' that may hamper student understanding of these ideas.

Each concept chapter is divided into five or six sections:

Section A. (Grouped together at the beginning of the book to facilitate workshop days.) Included here are laboratory activities (that in most cases can be modified for use as demonstrations). These activities may be used to introduce a topic by performing the activity and having students predict outcomes. This procedure is a useful way to elicit student pre-conceptions. Data sheets have been provided, and follow-up questions and alternative materials included where possible. These procedures may be used as they appear, or modified by additions or deletions to fit the time or content requirements of the different classroom situations or grade levels. These activities also form the basis of the laboratory activities for participants in STEM Resource Center programs.

Section B. This section includes background information and explication of the topic. Examples are given and important vocabulary words are underlined.

Section C. This section addresses common misunderstandings or misconceptions that students are known about these ideas. Since many of these erroneous notions are based on 'common sense' assessments of physical phenomenon or from misinterpretations of previous instruction, students (and teachers) have a hard time giving them up. It is very valuable to know what these ideas are before you teach a lesson on these topics since instruction can then be geared to elicit these wrong ideas and demonstrate or prove that they do not work. It has been found that students encounter difficulty in learning the real meaning (the implication of principles to real applications) of science ideas because they confuse 'ordinary' or everyday usage of words with the specific meaning of words as they are used in scientific definitions. These common word-confusions will also be pointed out in Section C.

Section D. This section contains a variety of suggestions for pre-instructional exercises, i.e., questions or problems to get kids started into a topic by stimulating their interest. Also included are knowledge-checking questions for use as follow-up exercises or for testing. A list of concepts is provided which may be used to construct concept maps of the larger target concepts.

Section E. A Glossary of Terms associated with the target concepts is provided.

Appendices of the Chemistry Resource Book contain essays on topics of interest to science teachers prepared by experts in various fields of science education. "What Research Has to Say to Science Teachers" rounds out some areas of important pedagogical knowledge useful for structuring classroom presentations. Also included is information concerning professional science teaching organizations, publications, and sources for obtaining information about field trips or software to enrich instruction, and science careers.

Section F. This section has selected web sites with information pertinent to the chapter content.

INTRODUCTION

Science Education: Critical Issues and Trends

Numerous studies show that most students make up their minds about science during their early school years. Judgments such as: science is hard; it's easy; it's dumb; it's boring; it's only for boys; it's fun; or I'd like to know more, are made by students based on what happens in these early experiences. What also has been established through educational research is:

What teachers do have effects on these decisions.

Our personal enthusiasm, the structure of our subject matter presentations, the depth of our own knowledge, our sensitivity and sense of fairness, the extent to which we make it all 'real' and 'relevant,' and communicate the expectation that <u>our students should change the way they view the world as a result of their experience with science</u>, have all been identified as pivotal factors in successful science teaching.

Success in a science course can be taken as achieving good grades. If everyone in an English class received A's, but never again read a book, that class was not successful. If everyone had perfect scores on a drug education test, and the used drugs – well – that's obvious. But we are finding that when students, even those who may have done very well on science tests, are asked to explain some aspect of the physical world, they do not use the information they were taught in school. They rely for their answers on self generated theories based on their own personal observations. These theories are often in direct contradiction to what was taught in science class. What can be the justification for the huge expense of education, if what is learned in school, doesn't transfer to students' lives when school is over?

Classrooms come in many different sizes and shapes. They may be homogeneously or heterogeneously grouped. Often they are overcrowded. But regardless of the particulars, all must be places where learning occurs. And this learning must be meaningful, that is, it must be available for retrieval when needed. What students learn in school should make a qualitative difference in the structure of their knowledge. New areas of science education research show great promise for helping us organize classroom practices to maximize meaningful learning. Findings from the area of cognitive psychology with its emphasis on understanding the process of human learning are being used to inform the art and science of teaching.

Structuring teaching based on how people learn, is a rather exciting trend. Instead of floundering around trying innovations on a hit or miss basis, we can be guided by the basic underlying mechanisms of our ultimate enterprise, i.e., learning. A complete understanding of human cognition is still in the future, but some important features of the process are beginning to emerge.

Learning used to be viewed as a 'filling up' with knowledge. We used terms like, 'getting it into their heads.' Given this view, it logically follows that teaching is a passing on of information:

teacher tells = student learns
which is
the "sage on the stage" model

However, if learning is understood in what is now referred to as a **'constructivist view,'** in which the learning actively puts together, or integrates new information with prior knowledge, a different approach to teaching is mandated. New information, that is, the subject matter we wish to teach, is not merely added to old information which the student possesses, but rather both the old and the new information are changed as a result of their interaction. Everything we already know, affects what is subsequently learned. We, as teachers, are not the central characters in our students' learning, they are. Learning is an internal change - it cannot be brought about solely from the outside. Does this perspective diminish the role and value of the classroom teacher? On the contrary, we are the producers, directors, supporters, cheerleaders, and expert guides, who set the stage and write the scripts for the learning process. What we say and do and how we say and do it, can assist or retard the mental processing our students must undergo. We become the "guide on the side". But it is the student who must undergo the process. What we say and do has to engage the student in processing information so that it fits in with his previous views about the world. In addition, we need to nurture inquiry and help students develop skills that allow them collect and process information and make good conclusions.

When we wish to teach a concept of modern science, such as, "matter is made of tiny particles called atoms," we are faced with the same problem that scientists have when they must convince the scientific community that they have new information to add to the existing body of knowledge. Only in our case, students also need to be taught the criteria for acceptance. We are really asking our students to view the world around them differently because of what we are saying. Their personal experience of the physical world does not lead them to conclude that matter is made of atoms, and in fact, many of their own perceptions are contrary to what we are calling the 'facts.' We do not see tiny atoms. We see big chunks of things. We see liquids flowing. If we only tell students that this is the way things are, because we say so, without demonstrating the 'why' behind what we think, and without recognizing and confronting ideas that they have already fashioned for themselves, they will keep the ideas that make sense to them. Science concepts will be tucked away along with old notebooks after the semester is over.

Many studies in which students have been tested or interviewed before and after classroom instruction tell us that children create for themselves a 'personal science' and hold on to these ideas even when they are contradicted in the classroom. They invent explanations for physical phenomena that seem perfectly reasonable to them, and to many adults. Some of the misconceptions that have been identified so far include: matter is continuous in nature as opposed to made of particles; burning is a destruction of mater, burning is a creation of energy (despite being taught the laws of conservation); current flows from the source of electricity to a bulb where the current is used up, (this is seen as reasonable in spite of being taught about circuits, since students see 'one' wire going to a lamp); plants get their food from fertilizers instead of making all their own food (students persist in this idea despite being able to recite definitions for photosynthesis). More and more of these common persistent erroneous notions are being identified as research continues. If science learning is to be meaningful, then the contrast between student ideas and science concepts must be made explicit.

Cognitive theory has lead to the following suggested steps for teaching to overcome student misconceptions:

First, find out what explanation students already have so that they can be contrasted to the science explanation. This may be accomplished by asking them to predict what will happen before we perform an experiment or demonstration. It is important to require that students provide reasons for their predictions. The classroom atmosphere must be such that students feel free to express their opinions. Second, perform the experiment, carefully laying out how the process leads to the conclusion. Clearly state the correct science concept, and emphasize how it explains the observed phenomena. Third, return to each student prediction and have them explain, why it didn't hold up, and in what way them might now change their thinking. Fourth, provide additional examples of cases where the correct concept can be used to explain things. It could be pointed out that many of the student predictions are reasonable, and were thought to be the explanation not very long ago before careful scientific experiments were performed.

Strategies incorporating the steps just described, have been shown to produce more students who have correct concepts, than even the most expertly performed classroom activities which do not specifically address the problem of misconceptions. If misconceptions are to be overcome, and students are to interpret the happenings in the world around them by referring to what they've learned in school, they must be convinced that their prior notions are no longer useful, and that the science concept is. These steps mirror the process by which new scientific discoveries become incorporated into accepted scientific knowledge.

One very simple way of checking that students have engaged in meaningful learning is to ask questions that require students to interpret some common experience. Like it or not, students gear their studying to satisfy the demands of teacher-constructed tasks. We engineer the mental processing of our students by what we require them to do. Exclusively asking questions that can be answered by rote memorization of definitions assures a low level of mental work, and also lowers the likelihood that the information will be available for retrieval from long-term memory when a relevant problem or situation arises. Retrieval from memory is facilitated when information is initially stored with a rich network of connections. Memorization of facts without reflection on how the ideas they represent fit in with the rest of what we know, will result in these facts being lost to us, unless they are frequently practiced. (Such is the fate of phone numbers or names we no longer use). Attention must be paid not only to what we test but how we test it as the process of studying dictates the level and meaningfulness of what is learned

The 5 E's Approach

A more structured method for developing activities that help students construct meaningful connections and dispel misconceptions is the 5 E's approach. This is based on the learning cycle introduced by Atkins and Karplus in 1962. The 5 E's are: engage; explore; explain; elaborate (or extend); and, evaluate.

Engage – This stage is designed to pique student interest and elicit students' current understanding so that connections can be made between what they know and what they will be investigating.

Explore – This part allows students to be actively involved in manipulating materials (or watching the teacher do so) and making observations that address the target concept. Not only does this provide experience that may contradict what they think should happen but it provides all students with a common experiential base to promote concept discussion.

Explain – Once students have explored they can now discuss what they observed, put correct vocabulary with what they did and recorded, and reconcile their former ideas with what they think as a result of the exploration.

Elaborate – In this stage, students can apply their understanding of the target concept to new situations to reinforce correct understanding or uncover problems with their knowledge. This also provides chances to relate science concepts to real-world situations.

Evaluate – This stage should actually happen throughout each of the 4 previous parts. This stage should allow the students and teacher to check the veracity of student knowledge and the development of process skills (see Appendix ?? for information on process skills).

Concept Mapping

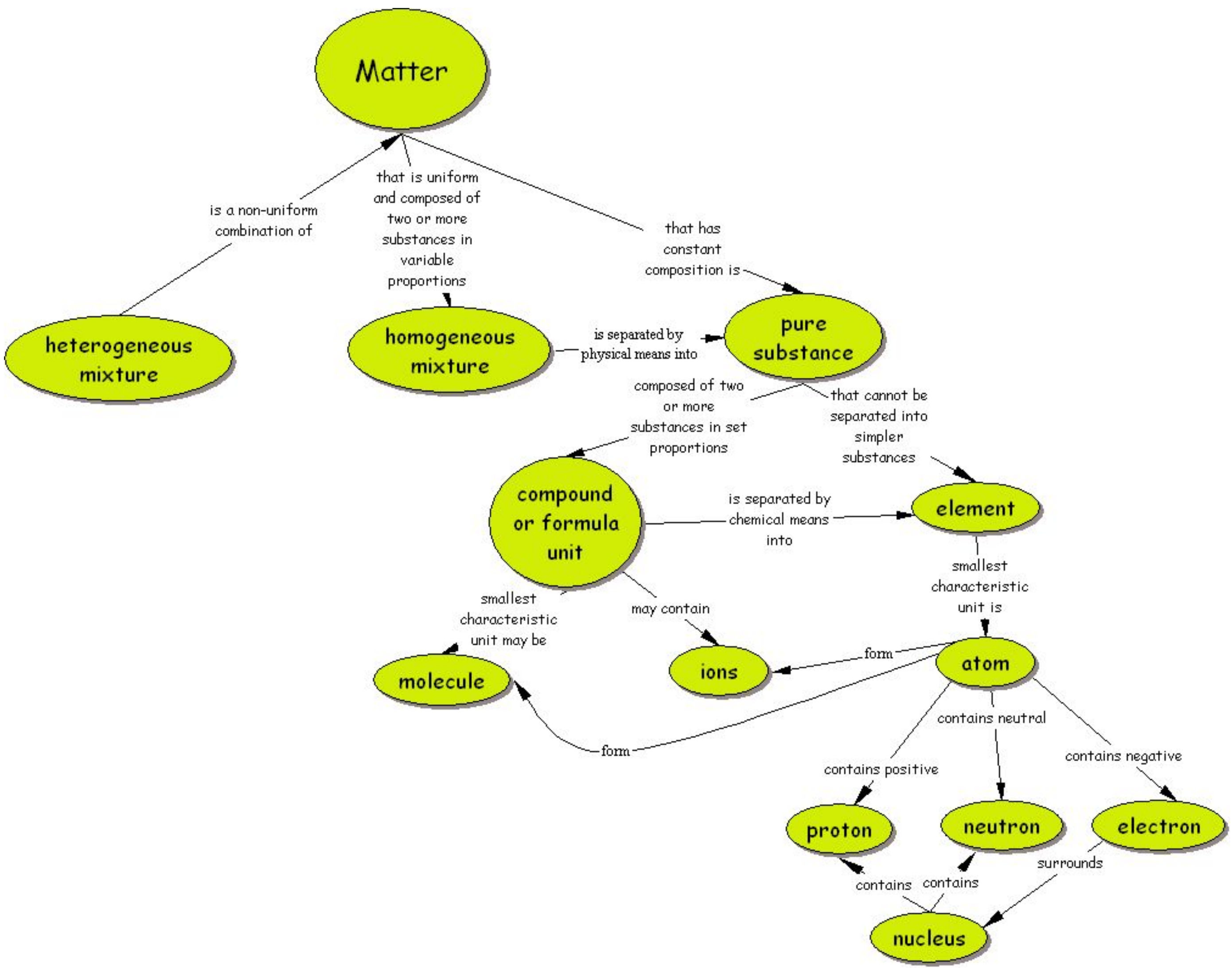
The constructivist view of learning theorizes that our knowledge is organized. New information has to fit in or be integrated with what we already know if we are to retain it and retrieve it when needed. The technique of concept mapping (fully discussed in and Novak Gowin's, Learning How to Learn, 1984), in which concepts are graphically displayed, is a means of helping student organize concepts to aid in "meaningful learning". A concept is defined as a regularity in events or objects designated by some label. For example, "restaurant" is a concept that covers many different establishments that have a property in common – serving food.

A concept map is not a simple list of important terms to be memorized. The concepts are enclosed in boxes or circles and arranged hierarchically with the broadest ones on top and more specific ones further down. Concepts that are closely related are placed at the same level. Arrows are draw between the concepts and words or phrases that succinctly describe their relationship are written over the arrows. The more connections that can be made, the richer that person's understanding of the concepts.

The construction of concept maps can be done by individuals or large or small groups. It is an excellent way to summarize the material in a unit, or show how two different units of material are related. It is also a useful way for teachers to plan instruction because it enables us to organize the curriculum so that we begin with the big ideas and then tie the supporting concepts back to the overall picture.

Concept mapping is easily taught to students. It is currently being used in many elementary and middle schools and by textbook publishers in their ancillary materials. When starting out to make a map, it may be a good idea to assist students in selecting the concepts they have learned so that the maps are limited to a few concepts and are not overwhelming. As students gain experience with the technique, they can select the concepts to include. It is best to keep the connecting words simple to begin with. Again, as students become more adept they can expand the connecting phrases. Although there can be incorrect aspects of student generated maps, such as wrong hierarchies or incorrect link terms, there can be many correct versions allowing individual creativity.

Following a unit on matter conducted with teachers, a map like the following could be constructed with the terms shown below.



Science Education Resources

PGCPS site providing links for alternate teaching strategies for science
http://science.uniserve.edu.au/school/support/strategy.html

School Improvement in MD web site with lots of curriculum resources for science
http://mdk12.org/instruction/curriculum/science/resources_other.html

One site of the National Association for Research in Science Teaching (NARST) that has articles applying research to the classroom
http://www.narst.org/publications/research.cfm

Maryland Voluntary State Curriculum (VSC) for science organized in an easy to use manner
http://mdk12.org/instruction/curriculum/science/index.html

CHAPTER 1

Matter: Properties, Changes, and Kinds

B. Background

The science of Chemistry studies matter and the changes matter undergoes. Matter is anything that occupies space and has mass. It is the physical stuff of which all material things are made. Energy is the ability to do work and can interact with matter but has no mass or volume of its own. Matter and energy do have something have in common- the conservation principle. The Law of Conservation of Matter says matter cannot be created nor destroyed in a normal chemical reaction. According to the Law of Conservation of Energy, energy cannot be created nor destroyed during a normal chemical reaction.

However, matter can be changed into detectable amounts of energy under certain special conditions, ex. nuclear reactions, atom bombs, the core of the sun. You are probably familiar with the equation $E = mc^2$. In words, this equation is "energy equals matter times the speed of light squared". We can determine how much energy can be derived from a given amount of matter. During the vast number of usual chemical and physical changes, the total amount of measurable matter present before a change is equal to the amount present after the change. The same applies to energy. Particles of matter have been created from energy, but only in the extraordinary conditions available in high energy particle accelerators.

Characteristics of Matter: Measurement

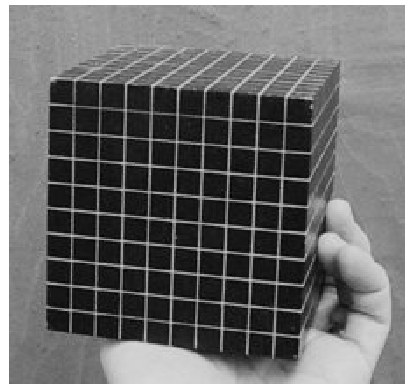
All matter has dimensions and the units used by scientists for describing those dimensions are SI units which are a modern Metric System. Keep in mind that measurement means comparing some aspect of matter to an accepted standard. This line, ___, is one centimeter long, not because there is anything one centimeter-ish about the line, but because this particular distance between two points is accepted around the world as being so. The Metric System offers the convenience of prefixes based on multiples of 10 which really makes dealing with quantities simpler than in our familiar English System. Some of the more common prefixes are in the table below.

Metric Prefix	Abbreviation	Value
kilo	k	1000 x
deci	d	1/10 or 0.1 x
centi	c	1/100 or 0.01 x
milli	m	1/1000 or 0.0001 x
micro	μ or mc	1/1,000,000 or 0.000001 x

If you want to make a unit 1000 times larger than a gram you add a “k” in front of grams. If you want a unit 1000 times smaller than a gram, add an “m” to gram.

1 kg = 1 kilogram	1 kg = 1000 g
1 mg = 1 milligram	1 mg = 0.001 g

(The kg is the SI unit for <u>mass</u> – the bulk of an object. Note that <u>weight</u> is a measure of the pull of gravity. We can interchange the two terms as long as we stay at sea level on Earth.)



A chemist must often specify the <u>volume</u> of a piece of matter or the space it occupies. The liter is the unit used for volume. If you have a square box that is 10 cm on each side, the space inside that box would equal 1 liter. The abbreviation for liter is “L”.

Volume cube = length x width x height
= 10 cm x 10 cm x 10 cm
Volume = 1000 cm^3 = 11Liter

1 milliliter, or 1 mL, is 1/1000 of a liter. 1 mL is the same as 1 cm^3.
The medical profession also uses the term 1 cc (cubic centimeter). 1 cc = 1 cm^3 = 1 mL
Scientist NEVER use cc and always use mL or cm^3.

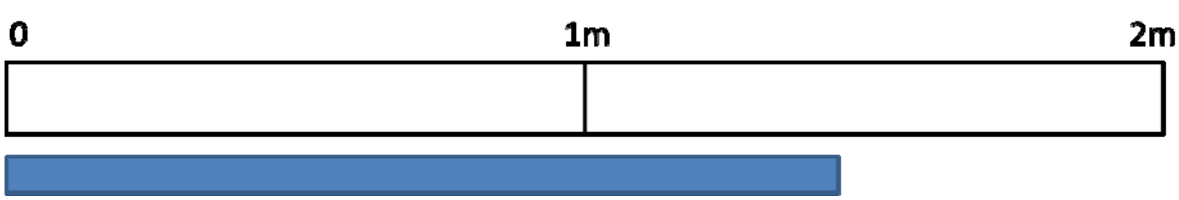
The meter is the unit for measuring distance. Common prefixes used for meter are:

1 km - 1 kilometer = 1000 meters
1 cm - 1 centimeter = 1/100^th^ meter = .01 meter
1 mm - 1 millimeter = 1/1,000^th^ meter = .001 meter

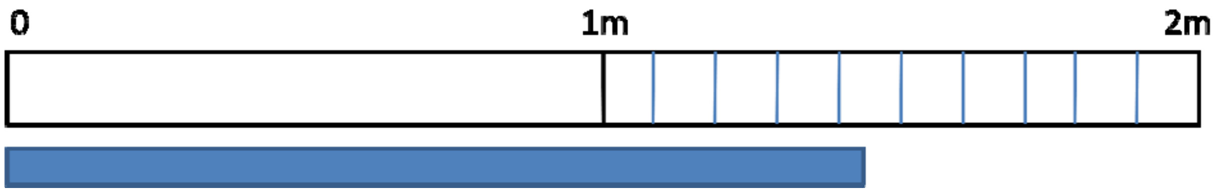
There are many other SI measurement units - the second for time and the ampere for electric current, for example. But kilogram, liter, and meter are fundamental to many basic chemistry concepts.

All measurements consist of three parts: (1) number; (2) unit; and (3) degree of uncertainty.

Care must be taken when measuring an object to use the correct number of <u>significant figures</u>. For example, if a piece of wood was measured along a measuring device that was marked only at the 1 and 2 meter points and its edge reached what appeared to be between the 1 and 2 meter mark, you would indicate its length as 1.5 m. The 0.5 is a guess. The significant figures in a measurement consist of all digits you are sure of and one that is a reasonable guess.



This last digit tells us to what extent we are uncertain. In this example, you know it is at least 1 m, and that it is not 2 m. That it is half-way is your guess so the degree of uncertainty occurs in the tenths place.



On a second measuring device (above) which is marked off in tenths of meters, an additional significant figure can be obtained. The measurement is 1.45 m, as you know for sure the length is greater than 1.4 but less than 1.5, so you make a guess that it between the two. The uncertainty occurs in the hundredths place. Therefore, the number of "sig figs" in a measurement is determined by the way in which the measuring device is marked off. Using the first device there were 2 sig figs in the measurement. Using the second one you could obtain 3 sig figs.

If a particular rectangular piece of wood were measured on it two sides by different measuring devices (side one = 1.5 m, side two = 1.45 m), and the area of the wood was calculated by multiplying one side by the other and the computation was done by calculator you will get:

Area = length x width
= 1.5 m x 1.45 m
Area = 2.175 m^2

The four significant figures in this answer would suggest that the measuring was done with a device that had more markings than any of those that were actually used. When multiplying or dividing measured numbers, the answer can have no more significant figures than the least amount in the problem. Therefore, the correct is 2.2 m^2, since the least number of significant figures was 2 (in 1.5 m). This rule applies when you are dividing measured numbers. When adding or subtracting measured numbers, the answer can have no more digits to the right of the decimal than the least amount in the problem. See the examples below.

6.239 cm
21.77 cm
100.8 cm
128.808 cm (by calculator)

CORRECT ANSWER = 128.8 cm

162.741 g
-94.6 g
68.141 g (by calculator)

CORRECT ANSWER = 68.1 g

Characteristics of Matter: Properties

In order to study matter, chemists must be able to describe it. Properties of matter can be categorized in two different ways, as **physical** and **chemical** properties or as **intrinsic** and **extrinsic** properties.

Physical Properties are descriptive of the matter itself without regard to how it reacts with other chemicals. An easy way of remembering the physical properties is by the acronym SCODS: State (solid, liquid, or gas); Color; Odor; Density (amount of matter in a given volume); Solubility (does it dissolve in specified solvents?). Also included are boiling and melting points, conductivity, viscosity, elasticity, malleability, ductility, and hardness.

Chemical Properties are descriptive of how matter reacts with other chemicals or how it behaves during chemical changes. For example, a characteristic might be whether something can burn. Burning is a chemical change in which a material rapidly combines with oxygen with the release of energy.

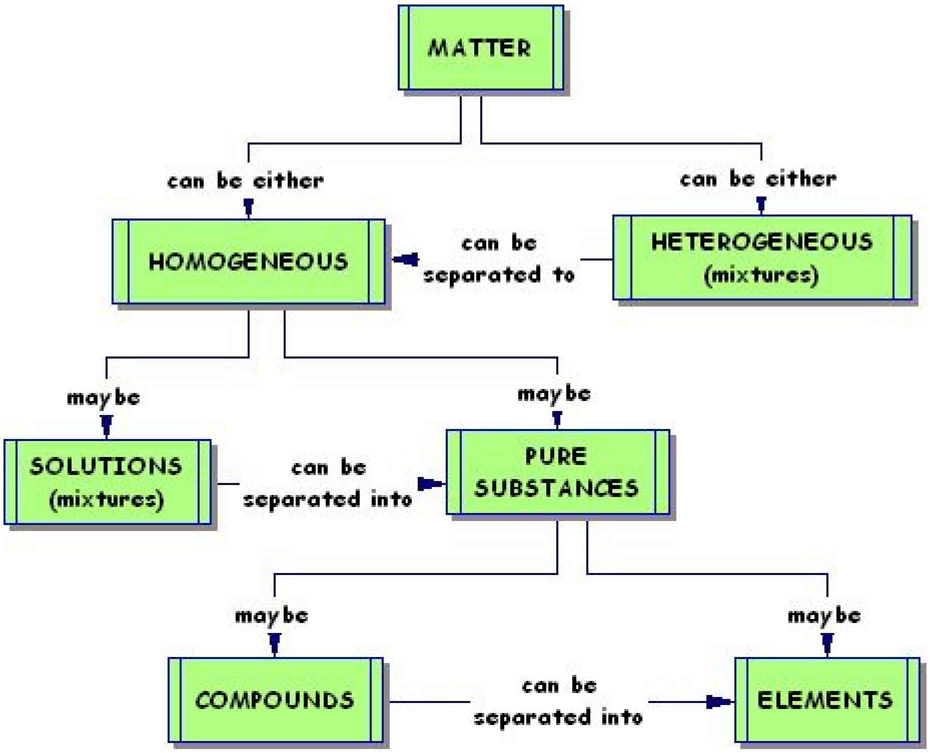
Intrinsic Properties are those characteristics that do not vary from sample to sample of one kind of matter, that is, they don't depend on the size of the sample. For example, sulfur is a yellow solid at room temperature no matter how much you have. Oxygen is a colorless gas. No matter how much sulfur or oxygen you have, they will have the same properties.

Extrinsic Properties are those characteristics that do change from depending on the amount of the same kind of matter. Examples include mass, volume or dimensions of the matter.

Changes in Matter: Physical and Chemical

When matter undergoes a physical change, the identity of the materials remains the same. It is the same "stuff" before and after although its appearance may have been altered. For example, air is a gas, but it can be made into a liquid if the temperature is lowered and the pressure increased. We can even make solid air. These changes in state are physical because the material is still air. We have not changed its chemical make-up. A chemical change however, alters the identity or chemical composition of a material. When hydrogen gas burns, it chemically combines with oxygen and forms a new material – water.

Classification of Matter: What Kinds of Matter are There?



Heterogeneous or non-uniform substances (mixtures) can be separated into homogeneous (even throughout) ones by sorting or filtration. Solutions (homogeneous mixtures) may be separated into pure substances by distillation or chromatography. Note that all of the processes to separate heterogeneous substances and solutions into pure substances are physical changes. The identity of the parts stay the same; they just become physically separated from one another. For example, when the solution salt water undergoes evaporation, salt is left behind and the water goes off into the air as water vapor. There was salt and water in the solution (just mixed together) and they remain salt and water after the process of evaporation.

Pure substances are of two types, elements and compounds. Elements are simple substances. There are 118 known elements, 92 of which are found in nature; the rest were made by man. Elements cannot be broken down in a chemical change. Compounds are composed of elements bonded together in a definite proportion by weight. There are many millions of compounds, some found naturally in the world, and many others synthesized or put together by chemists in the lab. They can be decomposed to elements during a chemical change.

C. Misconceptions

1. "Matter is lost in a chemical or physical change."

   When paper burns or iron rusts, we may perceive some matter as "disappearing." This is translated in many minds as lost. In reality, the matter has been transformed into a compound or element that is no longer visible to the naked eye. It is never lost. A student may think that water is lost when it is evaporated because he cannot see it. It is important for students to realize a form change does not mean a loss.

2. "Energy is lost during a chemical or physical change."

   The Law of Conservation of Energy (and matter) refers to conservation within a system. The largest system is the universe. A small system might be a closed, insulated container. Often, we perceive energy moving from a small system (reaction in a beaker) to a larger system (classroom) as "lost." It is not lost but may be distributed in a large region. It is important to remember that because it is not visible or felt, it is not necessary lost.

   Often chemical energy is <u>converted</u> to heat energy in the course of a chemical reaction. Heat "appears" from nowhere but has not been created, only released. Heat or light energy may also be stored as unseen chemical energy but is not destroyed.

3. The words "create" and "destroy" as they are used in the conservation laws have very specific meanings. Create – means to make out of nothing. Students often confuse this meaning with the ordinary acceptable usage of the word – in which we say, "I will create a new dress by sewing these pieces of material together." In this case, the idea of the new dress may be really new, but the physical stuff of the dress, the material was already there. Thus the dress is not a "creation" but an assemblage. When the Law says that matter cannot be <u>created</u> – it means arise out of nothingness.

   When we say that a house has been "destroyed" by a tornado, the physical parts are still in existence, albeit scattered apart. But when the Law refers to destroying matter or energy, it means that they have passed out of existence. Students have been known to confuse the scientific and ordinary meaning of these words and may refer to compounds as being destroyed when they are really just broken down into their parts.

D. <u>Warm-Up Exercises</u>

Before Lesson or Lab: Possible interest arousing questions

1. What is the world make of?

2. Why is it important that we know what things are made of?

3. Why do you think people in ancient times thought that everything was made of air, earth, fire, water? Did they see only these 4 things? How did they explain the existence of more than four things?

4. If matter has mass and takes up space, can you name anything that is <u>not</u> matter? How could you prove that something is or is not matter?

5. If the labels fell off jars of sugar and salt, how could you put the correct labels back on them? Do you think tasting chemicals in order to identify them is a good idea?

6. Make a list of some of the matter in this room. Is there some way we can classify it?

7. a. Describe properties of sugar (taste, color, state, solubility in H_2O).
   b. Describe what happens when you dissolve sugar in water. Are there changes? What are they? Is the sugar still there? How do you know that?
   c. What would happen if you burned sugar? What changes would you see? Is the sugar still there?

8. Decide if each of the following involves a physical or chemical change. Describe how you know which kind of change it is?
   a. Boiling water to steam
   b. Making Kool-Aid
   c. Scrambling an egg
   d. Cooking the scrambled egg
   e. Melting a popsicle
   f. Digesting a popsicle

<u>After the Activities</u>

1. Remember the properties of sugar we discussed.
   a. Were the changes when you dissolved sugar in water physical or chemical? Can you get the sugar back out of the water? How? What kind of change is that?
   b. Were the changes when sugar burned physical or chemical? Why?
   c. Is there anything else that behaves like sugar when is placed in water? Is there anything that behaves differently?
   d. Why do you think they call "elementary" school "elementary?"

2. Construct a concept map with the words: matter, solution, compound, element, homogeneous, heterogeneous, energy, physical change, chemical change.

3. Definitions of vocabulary terms. Give examples of each term. Use other words to describe these terms.

4. Perform the following calculations.
   Round off the answers to the correct number of significant figures.

   a. $4.95\text{ m} \times 3.625\text{ m} =$ Ans. 17.9 m^2

   b. $\dfrac{100.63\text{ kg}}{5.2\text{ kg}} =$ Ans. 19 kg

c.	6.3106 m	=	Ans. 41.9 m
	.57 m		
	32.1 m		
	+ 2.931 m		

E. Glossary

Absorb	take up one substance into the bulk of another substance
Adsorb	adhere to the surface
Boiling	a process recognized when rapid evaporation takes place below the surface of a liquid
Boiling point	temperature at which a liquid and gas are at equilibrium (liquid ↔ gas)
Chemical change	change that produces matter different from the original; a change in the identity of matter
Chemical property	property describing how matter will change in a chemical change
Chromatogram	resulting product of chromatographic separation; in paper chromatography, it is the piece of paper with the components located at various points from bottom to top along the paper.
Chromatography	separation technique based on different solubilities of solution components between moving and stationary media
Compound	pure substance made up of 2 or more elements in a fixed composition that can be broken down into these elements by chemical change
Conductivity	ability to conduct an electrical current
Density	mass per unit volume; it is a measure of the tightness of packing of particles.
Dissolve	evenly distribute solute in a solvent; a physical change
Distillation	separation technique where volatile (low boiling point) liquid is evaporated (converted to a gas) and then condensed (converted to a liquid) into a separate matter
Ductility	ability to be drawn into wire; an example of a ductile material is copper

Elasticity	ability to regain shape after being deformed
Element	pure substance that cannot be broken down into anything simpler during a chemical change
Equilibrium	a condition that exists when two opposing processes are taking place at a constant rate
Evaporation	changing liquid to gas; a physical change
Filtrate	liquid collected during filtration; it is what comes through the filter paper
Filtration	process of separating a liquid from a solid by pouring the mixture through filter paper
Heterogeneous	different throughout
Homogeneous	same throughout; if a sample is taken from any part of a homogeneous substance, it will be identical in identity and composition to any other part
Kilogram	standard unit for mass in the metric system
Law of Conservation of Energy	energy cannot be created nor destroyed; during a chemical or physical change
Law of Conservation of Matter	matter cannot be created nor destroyed; during a chemical or physical change
Malleability	ability to be hammered into a sheet; an example of a malleable material is gold
Mass	the bulk of an object
Matter	anything that occupies space
Mixture	a combination of two or more substances that are physically mixed, not chemically combined; mixtures can be heterogeneous or homogeneous
Moving phase	phase that moves in chromatography
Physical change	change that does not involve change in the composition of matter; the substance maintains its identity; although, it may look different

Physical property	characteristic of a substance that does not involve chemical change
Solubility	maximum amount of substance (solute) that will dissolve in another substance (solvent)
Solute	substance that is dissolved in making a solution
Solution	homogeneous mixture of two or more substances that has a variable composition
Solvent	substance that promotes dissolving in making a solution; when two liquids form a solution and one of them is water, the water is considered to be the solvent
Solvent front	point to which solvent rises in paper chromatography
Stationary phase	non-moving phase in chromatography, in paper chromatography is the paper
Viscosity	resistance to flow of a liquid; molasses has a higher viscosity than water
Weight	a measure of the pull of gravity on an object

F. Additional Resources

National Institute of Standards and Technology web site on SI units
http://physics.nist.gov/cuu/Units/introduction.html

A tutorial on significant figures
http://tourserver.rice.edu/documents/SignificantFigureRules1.pdf

CHAPTER 2

Atoms and Elements

B. Background

Elements and Atomic Structure

The majority opinion among current curriculum developers is that topics for elementary school science should be restricted to those concepts that can be concretely manipulated by children. Since atoms and molecules are too small to be seen, it is suggested that their detailed study not be included in K-6 classrooms. It is important, however, for teachers to understand these abstract concepts as they form the basis for explaining the behavior of matter on the scale that we can see.

An element is a substance that cannot be broken down into other substances in a chemical change. Gold, oxygen and sulfur are examples of elements. If you take the smallest unit of an element that has the fundamental characteristics of that element you have an atom. The atom is the smallest unit that can enter into a chemical combination. Each element or atom is given a one or two letter symbol called the atomic symbol. For two letter symbols, the first letter must be uppercase, and the second letter lowercase. For example, oxygen is O, calcium is Ca, and copper is Cu. Some symbols derive from Latin names and at first do not seem to fit. Kalium is the Latin name for the element potassium and the symbol is K.

Investigations in the mid-1800 to the mid-1900 revealed some basic information about the structure of the atom. It is known that atoms have parts. Each atom contains three major types of sub-atomic particles: protons, neutrons, and electrons. The table below provides information on these particles.

Table 2.1. Sub-atomic Particles

Particle	Charge	Mass (amu*)	Location
Proton	1+	1.0073	nucleus
Neutron	0	1.0087	nucleus
Electron	1-	0.00055	outside nucleus

---

*a.m.u. stands for atomic mass unit. It is impossible to actually weigh these tiny bits of matter, so instead their masses are compared to a single atom of C which is assigned a mass of 12 a.m.u's. Therefore, a proton is about 1/12th the mass of a carbon atom. For all practical purposes, electrons contribute no mass to an atom. It is now known that 1 a.m.u. = 1.660 x 10^{-24} g. (If you realize that there are about 454 grams in 1 pound, and move the decimal 24 places to the left to write 1.660 x 10^{-24} in standard form, you can get some idea of the incredibly small size of atoms and their particles.)

Each element varies in the number of protons its atoms have. The number of protons is the atomic number. This number can be found on the Periodic Table. It is the smaller of the two numbers given along with the symbol and generally written in the top of the element box. Since atoms are electrically neutral, the number of protons and electrons in each atom is equal. However, the numbers of neutrons vary, even for atoms of the same type. The number of protons plus neutrons is termed by the atomic mass number. It is a whole number and is not on the Periodic Table. Atoms of carbon (C) must have 6 protons and 6 electrons but may have 6, 7, or 8 neutrons. When two atoms have the same atomic number but different mass numbers (because of having a different number of neutrons), they are called isotopes. Therefore, we refer to C-12, C-13, C-14, as the three isotopes of carbon. They are all C atoms, the fact that they have 6 protons determines this, but they vary in mass.

6
C
12.011

mass number = number of protons + number of neutrons

We can also write the symbols for the carbon isotopes as see below where the superscript is the mass number and the subscript is the atomic number. For these formulas the super and subscripts are written to the left of the atomic symbol.

$$ {}^{12}_{6}C = C\text{-}12 \qquad {}^{13}_{6}C = C\text{-}13 \qquad {}^{14}_{6}C = C\text{-}14 $$

The larger number associated with each element on the Periodic Table is called the atomic mass. It is determined by taking the mass, as compared to an atom of C-12, of each isotope of that element and averaging them according to the percentage of that isotope found in nature.

If you want to know the number of protons or electrons that an atom of a particular element has, it is the same as the atomic number. If you want to know how many neutrons an atom has, you would have to know which isotope you are referring to and subtract the atomic number from the atomic mass number. If you round-off the atomic mass on the Periodic Table and subtract the atomic number, you obtain the number of neutrons in an average atom of that element.

The nucleus is the dense central region of the atom and contains the protons and neutrons. According to one model (Bohr model) for atomic structure; the electrons move in defined orbits around the nucleus. Electrons can only move in these orbits called shells and the electrons in each shell have a definite amount of energy. This idea came from studies of “excited” atoms which showed that electrons absorbed only certain quantities of energy to become excited, emitting the same energy when they returned to ground state from the excited state. The exact value for the energy absorbed or emitted depends on the atom (#p, #e). Atoms of each element are unique. An atom’s electrons may absorb heat energy but will release the energy as visible or ultraviolet light. So when we place various elements in the flame of a Bunsen burner, the flame will turn various colors depending on the element. The color of light is determined by its wavelength and its energy. Using a spectroscope, one can see the individual quantities of light (colors) and use this to identify the element.

Studies in the middle of this century modified the solar system atomic model. The Bohr model has been changed. We no longer picture electrons travelling around the nucleus in circles the way planets revolve around the sun. Instead electrons are viewed as existing in the obitals – a probable region in space within which 2 electrons move around the nucleus. The orbitals differ in shape. "s" orbitals, for example, are spherical, and each shell or energy level starts off with one "s" orbital. "p" orbitals are dumbbell shaped. Starting with the second shell, each shell has 3 "p" orbitals. The electron can move from region to region but cannot reside between. Each shell contains from 1 to 16 orbitals.

The first shell or energy level (K) contains only 2 electrons, both of them in an "s" orbital. The second energy level (L) can contain a maximum of 8 electrons, 2 "s" electrons, and 6 "p" electrons (2 in each of 3 "p" orbitals). The third shell (M) can contain a maximum of 18 electrons, 2 "s" electrons, 6 "p" electrons, and 10 "d" electrons (2 in each of 5 "d" orbitals).

Element Characteristics and the Periodic Table

All known elements have been organized on a chart called the Periodic Table. Each element occupies a box which contains at least two numbers. One whole number is the atomic number. The other number is a decimal value called the atomic mass (in a.m.u.). The atomic mass is the average mass of all isotopes of the element and is relative to an isotope of carbon ${}^{12}_{6}C$. The Table is set up so that atomic number increases as you move from left to right. Elements with similar chemical properties are placed in a column called a family or group. Elements in a row occupy the same period. The period numbers (Arabic numbers from 1 to 7) tell how many shells or energy levels the atoms of that element contain. The Table is divided by a "stair-step" line into two unequal sections. Elements to the left of the stair step are metals. Elements to the right are non-metals.

As you move down a group, the sizes of the atoms increase. This is because each successive atom has an additional shell of electrons. In general, as you move across a period the size of the atoms decreases because electrons are being placed in the same shell (or a lower shell) and the extra protons added to the nucleus cause the electrons to be drawn in closer to the nucleus.

Periodic Table of the Elements

Chemistry Reference Sheet

California Standards Test

Key

11 — Atomic number Na — Element symbol Sodium — Element name 22.99 — Average atomic mass*

	1 1A	2 2A	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
1	1 **H** Hydrogen 1.01																	2 **He** Helium 4.00
2	3 **Li** Lithium 6.94	4 **Be** Beryllium 9.01											5 **B** Boron 10.81	6 **C** Carbon 12.01	7 **N** Nitrogen 14.01	8 **O** Oxygen 16.00	9 **F** Fluorine 19.00	10 **Ne** Neon 20.18
3	11 **Na** Sodium 22.99	12 **Mg** Magnesium 24.31											13 **Al** Aluminum 26.98	14 **Si** Silicon 28.09	15 **P** Phosphorus 30.97	16 **S** Sulfur 32.07	17 **Cl** Chlorine 35.45	18 **Ar** Argon 39.95
4	19 **K** Potassium 39.10	20 **Ca** Calcium 40.08	21 **Sc** Scandium 44.96	22 **Ti** Titanium 47.87	23 **V** Vanadium 50.94	24 **Cr** Chromium 52.00	25 **Mn** Manganese 54.94	26 **Fe** Iron 55.85	27 **Co** Cobalt 58.93	28 **Ni** Nickel 58.69	29 **Cu** Copper 63.55	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.72	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.92	34 **Se** Selenium 78.96	35 **Br** Bromine 79.90	36 **Kr** Krypton 83.80
5	37 **Rb** Rubidium 85.47	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.91	40 **Zr** Zirconium 91.22	41 **Nb** Niobium 92.91	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.91	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.87	48 **Cd** Cadmium 112.41	49 **In** Indium 114.82	50 **Sn** Tin 118.71	51 **Sb** Antimony 121.76	52 **Te** Tellurium 127.60	53 **I** Iodine 126.90	54 **Xe** Xenon 131.29
6	55 **Cs** Cesium 132.91	56 **Ba** Barium 137.33	57 **La** Lanthanum 138.91	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.95	74 **W** Tungsten 183.84	75 **Re** Rhenium 186.21	76 **Os** Osmium 190.23	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.97	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.38	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.98	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
7	87 **Fr** Francium (223)	88 **Ra** Radium (226)	89 **Ac** Actinium (227)	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (266)	107 **Bh** Bohrium (264)	108 **Hs** Hassium (269)	109 **Mt** Meitnerium (268)									

58 **Ce** Cerium 140.12	59 **Pr** Praseodymium 140.91	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.96	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.93	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.93	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.93	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.97
90 **Th** Thorium 232.04	91 **Pa** Protactinium 231.04	92 **U** Uranium 238.03	93 **Np** Neptunium (237)	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (262)

* If this number is in parentheses, then it refers to the atomic mass of the most stable isotope.

Dealing with Numbers and Masses of Atoms: The Mole

Individual atoms are incredibly small. As a result we generally cannot deal with small quantities of them. We cannot place 10 or even 1,000,000 atoms in a test tube, as it is necessary to remove 500,000,000,000,000,000 (5×10^{17}) copper atoms from a penny before you could even detect a change in weight of the penny using the most sensitive scale on earth!

The way out of this problem is to use a standard for measuring quantities of atoms – (an SI unit of measurement) called the MOLE. The idea of a mole is similar to that of the DOZEN. A dozen means 12 units of anything. A mole means 6.02×10^{23} things. So if we have a mole of a particular element it contains 6.02×10^{23} atoms of that element (that's 602 followed by 21 zeros). This is also referred to as Avogardro's Number. Now how can we know when we have that many atoms? If you mass out on a balance the atomic mass of an element (from the Periodic Table) in grams, that mass is the molar mass or the mass that contains the Avogadro number of the atoms. For example, 24.31 g of magnesium contains 6.02×10^{23} magnesium atoms but it only takes 12.011 g of carbon to provide 6.02×10^{23} atoms of carbon. The different in the molar masses reflects the different masses of the individual atoms. You can see that carbon atoms are about half as massive as magnesium atoms.

We can also apply the mole to help us deal with quantities of compounds. For example, the smallest individual piece of the compound H_2SO_4 (sulfuric acid) is a molecule. If we add up the molar masses of each of the elements in this compound, we will have the mass of a mole of H_2SO_4 and that mass will contain 6.02×10^{23} molecules of H_2SO_4. Example:

Element	Molar Mass	# of Atoms in the Molecule		Total Weight
H	1.0079	x 2	=	2.0158
S	32.06	x 1	=	32.06
O	15.9994	x 4	=	63.9976
				98.0734

98.07 g = 1 mole of H_2SO_4

If you had only 49.04 g of H_2SO_4, that would equal 0.5 moles and would contain ½ (6.02×10^{23}) molecules of H_2SO_4. You can calculate the exact number of molecules of a compound or atoms of an element in a particular sample as long as you know the mass of one mole.

$$\#\ \text{moles} = \frac{\text{grams}}{\text{molar mass (mass of one mole)}}$$

$$\#\ \text{particles} = \#\ \text{moles} \times 6.02 \times 10^{23}$$

The mole is also a useful way of describing how much of a substance is dissolved in water. If a bottle is marked 1 M H_2SO_4, it means that enough water was added to 98.07 g of H_2SO_4 (1 mole) so that it equals a volume of 1 liter. This is referred to as a 1 molar solution of H_2SO_4. It is therefore possible to "pour out" a given number of H_2SO_4 molecules because we know exactly how many molecules are evenly distributed in the 1 liter of sulfuric acid solution. How many molecules would you have in 250 mL of a 1 M solution?

C. Misconceptions

1. "Between the nucleus and electrons in an atom there is air."

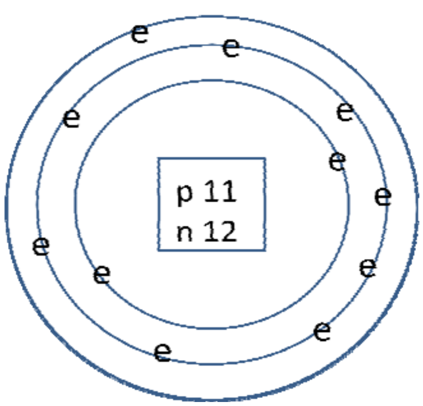
Many people think that air must be in the empty space between the nucleus and electrons of atoms or molecular. When we are young we are taught that even though we cannot see the medium surrounding us it does exist. So we grow up believing that any empty space must be occupied by air. But in fact – there is NOTHING in the space between the nucleus and electrons. Incidentally, relatively speaking, there is a vast region of "emptiness" or "nothingness" in atoms. If you imagine increasing the size of a nucleus until it was the size of a peanut, the first electron would be about a half a mile away. Atoms are mostly empty space.

2. A single atom of an element will have the same physical properties as the element.

Many physical properties such as conductivity, luster and boiling point are properties that a collection of atoms show. The fact that metals are shiny is based on the way light is reflected from a metallic crystal. Conductivity in a metal occurs because electrons move freely from one atom to another. A collection of atoms will, on average, boil at a particular temperature. However, a single atom would already be a gas, as gases consist of particles that are far apart from one another. It is incorrect to assume that a single atom will show the same characteristics as the element.

3. Atomic models may generate problems.

When we use models for atoms, we may inadvertently convey misconceptions to students. All models stand for something else – but they are not exact replications or pictures. For example, if we drew on the blackboard the following atom diagram representing a single atom of sodium – what wrong ideas might it convey?



1) that nuclei are square and have the protons on one end and neutrons on the other.
2) that electrons travel individually around the nucleus in circular paths.
3) that the distance between shells is about the same.
4) that the first shell is quite close to the nucleus
5) that individual atoms can be seen.
6) electrons and atoms are perfectly still.
7) atoms are flat.

Atomic models are still useful however, and pointing out that they accurately represent only certain aspects of atoms (in this case the number of particles, number of electrons in each shell, and division between particles inside and outside the nucleus) and that other aspects are misrepresented can be an added bonus to learning.

4. Troublesome words in the context of atoms and elements are:

   1) Shells – Students associate this word with eggs and visualize a shell as a hard surface. It must be pointed out that atomic shells are energy levels – or 3-dimensional places in space occupied by electrons of similar energy. It is useful to think of "orbitals" as clouds of electrons. If one atom approaches another atom, these negative electron clouds can actually become distorted in shape. When the size of atoms is measured (by means of X-rays studies) – it must be specified what environment the atoms was measured in – because what is around an atom can influence its radius.

   2) Empty space – means a vacuum, or the total absence of matter. Human beings have trouble with this notion. It is difficult to conceptualize "nothingness."

D. Warm-Up Exercises

Focusing Questions

What is the smallest thing we can see?
What instruments help us to see small things?
Can atoms be seen under a microscope?
Do you believe that there are such things as atoms? What evidence do we have to go on?
If you have a paper bag and you couldn't see through the bag, what could you do to help you figure out what was in the bag without opening it and looking inside?

Before the Activity

1. Thinking about the tiniest thing you can imagine. Could you break it down any more? What would you get?

2. Look around for as many shiny things as possible. Why are they shiny? Do they have any other properties in common?

After Lesson or Lab

1. Make a concept map with the terms: atoms, electron, atomic number, isotope, proton, neutron, nucleus.

2. Explain:

    1) How an atom of one element is different from another
    2) How two atoms of the same element may differ from each other
    3) Where each part of an atom is
    4) Why atoms are arranged the way they are on the Periodic Table
    5) Draw simple atom diagrams of the first 10 elements on the Periodic Table
    6) Select an unfamiliar element. What can you know about this element based only on its location of the Periodic Table

3. Build an atom by:

    1) Using small Styrofoam balls and toothpicks; balls should be different colors and larger for protons and neutrons

    2) Using small colored circles cut from three colors or paper, assemble the atom on a sheet of paper

4. Define vocabulary terms. Give examples of each term. Rephrase the definition in words different from those used in the textbook.

5. Work practical problems using the mole concept

    1) How many bits of NaCl are in a 1-pound (l lb) box of salt? (Remember to convert 1 lb to g. Divide g by the mass of one mole of NaCl and multiply by Avogadro's number).
    This problem can be made easier by asking it in steps:
        a. How many g are in a 1 lb box of salt?
        b. What is the mass of one mole of salt?
        c. How many moles are in a 1 lb box?
        d. How many particles of NaCl are in one mole?
        e. How many particles of NaCl are in a l lb box?

    2) If you had 6.02×10^{23} atoms of neon, what is its mass?

    3) What is the mass of one atom of calcium?

6. Write a biography of an element! Pretend the element is a person - where does it live - who are its relatives; personality characteristics; what kind of work does it do? (etc., etc.) Students can read their biographies out loud and have the class guess which element they are describing. The "stories" can be compiled and distributed to all the students. This is a fun way of combining writing, library work, and science, and offers opportunities for creativity in expression.

E. Glossary

Atom	smallest characteristic part of an element; smallest unit that can enter a chemical reaction
Atomic Mass	average mass of naturally occurring isotopes of an element relative to ^{12}C
Atomic number	number of protons in atomic nucleus; unique for each element
Atomic symbol	one or two letter abbreviations for atomic name (ex.: Ca is a symbol for calcium)
Avogadro's number	number of objects in one mole of a substance; 6.02×10^{23} objects
Electron	negatively charged atomic particle, very small and located outside the nucleus
Excited state	state of an electron (or molecule or ion) higher in energy than normal (ground) state – the electron has been promoted by the addition of energy to a higher energy level
Family	column of elements on Periodic Table with similar chemical properties; also called a group
Flame test	test to identify elements based on the color they exhibit in a flame
Ground state	lowest (normal) energy state of an electron (or molecule or ion)
Group	column of elements on the Periodic Table with similar chemical properties; also called a family
Isotope	atom of the same element with the same number of protons but different number of neutrons
Mass number	sum of the protons and neutrons in an atom
Metal	substance with characteristics of luster, malleability, conductivity; located on left side of stair-step demarcation of the Periodic Table

Model	representation of system or object
Molar mass	mass of Avogadro's number of a substance; the atomic or molecular mass expressed in grams
Mole	a quantity of anything that contains 6.02×10^{23} individual pieces
Non-metal	element that does not show metallic properties; located to the right of the stair-step demarcation of the Periodic Table
Neutron	neutral atomic particle; same size as proton and located in the nucleus
Nucleus	dense central region of an atom; composed of protons and neutrons
Orbital	probable region in space where an electron can be found; these vary in shape and contain a maximum of two electrons
Period	row of elements on the Periodic Table; all the elements in the same period have the same number of energy levels
Periodic Table	arrangement of elements based on atomic number and chemical properties
Proton	positively charged atomic particle; same size as a neutron and located in the nucleus
Spectrum (spectra)	pattern of light wavelengths characteristic of an element

F. Additional Resources

Good source of printable periodic tables of many varieties
http://www.sciencegeek.net/tables/tables.shtml

CHAPTER 3

Molecules, Ions and Bonding

B. Background

As a result of many experiments in which chemicals were carefully measured before and after changes, John Dalton formulated the Atomic Theory (1808). Part of this theory states that compounds form when two or more elements combine. He also noted that a particular compound always contains the same elements in the same proportions by mass. The second statement is actually the Law of Constant Composition or Definte Proportions. Work with his atomic theory led him to postulate the Law of Multiple Proportions. It says that two elements may form more than one compound but the masses of each element in the compound are in a ratio of small whole numbers.

These ratios and constant compositions are exemplified in the formula for compounds; CO, carbon monoxide, that has one carbon and one oxygen atom; and CO_2, carbon dioxide, with one carbon and two oxygen atoms. They are distinct compounds with different properties, that contain the same elements, but the proportion of the atoms of each element is different.

The smallest unit of a compound with composition of the compound is a molecule. The atoms within a molecule are held together by a force called a bond. A formula such as K_2CO_3 indicates that 2 atoms of potassium, 1 atom of carbon, and 3 atoms of oxygen are bonded together as one unit. For $A1_2(SO_4)_3$, 2 aluminum atoms, 3 sulfur atoms and 12 oxygen atoms are linked. The number following the atomic symbol is called a subscript. The subscript refers only to the element it follows or the elements within parentheses. If we write, $2K_2CO_3$, it means there are 2 molecules of the compound K_2CO_3. The 2 is a coefficient, i.e., a number written in front of a formula that tells the number of molecules.

NOTE: There are certain compounds, i.e., those that are held together by ionic bonding, that do not have molecules as their smallest unit. Some compounds, as will be discussed later, form an array of oppositely charged ions. The smallest unit of ionic compounds is referred to as a formula unit.

Why Do Atoms Combine to Form Compounds?

It was pointed out in the earlier chapter on atoms, that when energy is added to electrons they move up to a higher energy level, but they quickly fall back and release the additional energy. It is generally observed in nature, that matter is most stable when it is in a condition of low energy. Hot objects cool off all by themselves, a rock that is perched on the edge of a hill (it possess high potential energy due to its position) requires little inducement to tumble down the hill (where it has less potential energy). The making of bonds is an energy releasing process. By forming a bond between them, atoms possess less energy than they have as individual atoms and so bond making makes them more stable.

It is known that a group of elements are inert – or do not tend to form bonds with other elements (He, Ne, Xe, and Rn). These elements all have 8 electrons in their outer shell or energy level, (in the case of He, the outer shell has 2 electrons which is the maximum the first energy level can accommodate). The condition of having 8 electrons on the outer shell appears to be a stable one for atoms, and atoms that do not have 8 electrons to start with, react with other atoms either by transferring electrons among themselves (IONIC BONDING) or sharing electrons (COVALENT BONDING) until an "OCTET" of electrons in all the atoms is achieved. This tendency to achieve an octet of electrons is known as the OCTET RULE. But achieving 8 electrons is not why atoms combine; atoms combine in order to become stable. There are many examples of compounds in which the octet rule is broken. Note that the nucleus of an atom does not change during bond formation. It is also true that if bond formation or making is an energy releasing process, then bond breaking requires energy.

The group number from the Periodic Table indicates the number of electrons on the outer shell of the "A" group elements. Magnesium (Mg) is in group IIA and has 2 electrons on its outer shell, as does Ca and Ba. Elements in group VIIA (F, Cl, Br, I) have 7 electrons in outer shells.

Elements with 5 or more electrons (non-metals on the right side of the Table) will usually gain electrons to achieve an octet. Here is an example using oxygen:

$$O + 2e^- \text{---->} O^{2-}$$

$$6e^- + 2e^- = 8e^-$$

$$6p^+ = 6p^+$$

Because oxygen now has more electrons than protons it is negatively charged. Atoms that have a charge are called ions. Negative ions are called anions. Nitrogen which is in Group 5 with form N^{3-} and chlorine in Group 7 will form Cl^{1-}.

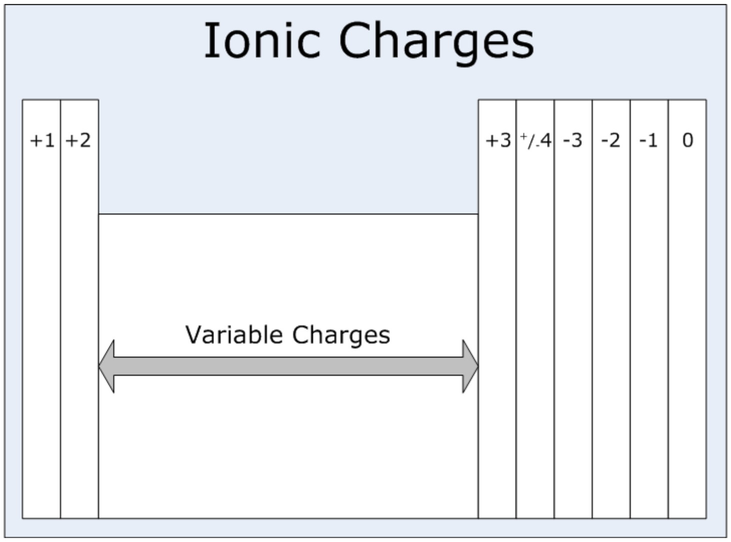
Elements with 3 or less electrons (metals on the left side of the table) will give up electrons because this will then expose the next inside shell which has 8 (or 2 if small) electrons.

$$Mg - 2e^- \text{---->} Mg^{2+}$$

$$12e - 2e^- = 10e^-$$

$$12p+ = 12p^+$$

These elements will have a positive charge since the number of protons is greater than electrons. These positive ions are called cations.



Once cations and anions form they are attracted to each other because they are oppositely charged. This attraction is called an <u>electrostatic</u> <u>force</u>. The attraction is strong enough to keep the ions in rigid formations in pure form. The strong attraction is called an <u>ionic</u> <u>bond</u> and the resulting product is an <u>ionic</u> <u>compound</u>. The formula depends on the charges of the ions:

Na^+ + $2e^-$ ----> $NaCl$

Mg^{2+} + $2Br^-$ ----> $MgBr_2$

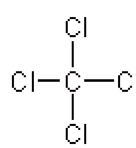
K^+ + S^{2-} ----> K_2S

Since compounds are neutral or have no charge, the positive and negative charges must cancel each other out. In the case of $MgBr_2$, it takes two bromide ions to accommodate the two electrons that magnesium lost. Each bromine can accept only one electron, because the atom already has seven electrons of its own.

Ionic bonds generally form when elements with less than 3 electrons on outer shells are combined with elements that have 5 or more electrons – or between metals and non-metals.

Elements with 4 or more electrons may also combine by sharing pairs of electrons. The electrons are shared by the nuclei of both atoms.

Cl + Cl ----> Cl : Cl

C + 4 Cl ----> 

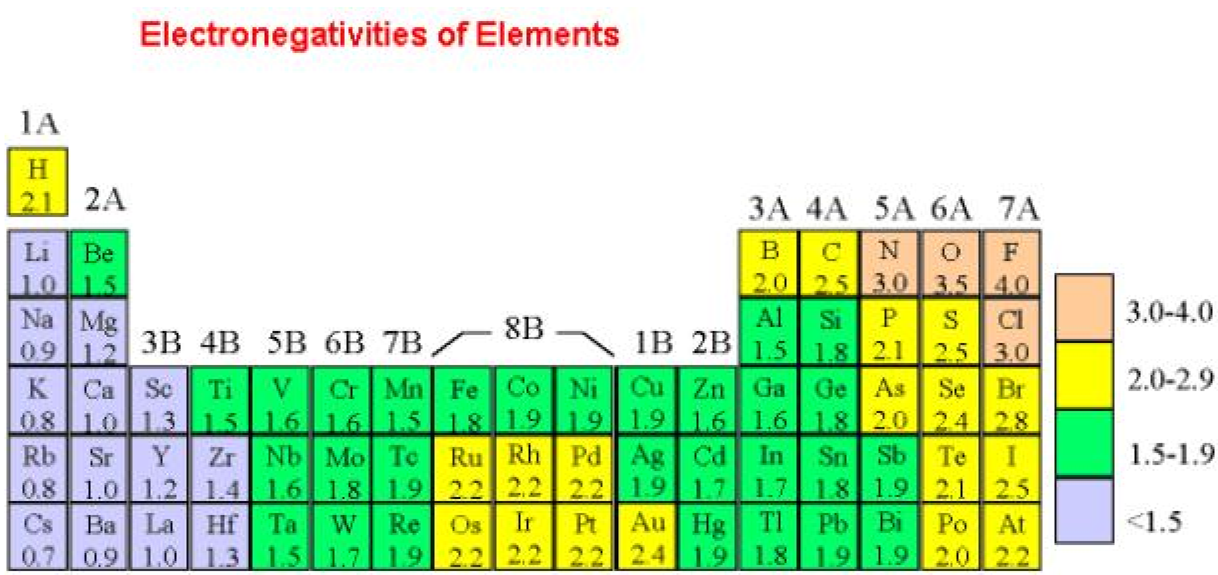
$C1_2$ and $CC1_4$ are considered to be <u>covalent</u> compounds because the molecules were formed by the sharing of electrons. The atoms have to stay near each other so the electron clouds can overlap. The diagram used above for Cl_2 is a <u>Lewis Dot Structure</u>. These diagrams consist of the symbol for the element and dots representing elements in the <u>valence</u> <u>electrons</u> or the electrons in the outermost shell rather than using a line to represent the shared pair of electrons.

Atoms may share one, two, or three pair of electrons:

C1 : C1	O :: O	N ::: N
1 pair shared Single bond	2 pairs shared double bond	3 pairs shared triple bond
Each chlorine atom has 7 electrons, by sharing one electron between them, each atom now has 8 electrons.	Each oxygen has six electrons, and by sharing 2 electrons they each now have 8 electrons.	Each nitrogen begins with 5 electrons, so it is necessary to share 3 electrons to achieve an octet.

The sharing of electrons creates a bond called a covalent bond. The resulting compound is a covalent compound. These compounds are generally formed by two or more non-metals.

The sharing of electrons in a covalent compound may or may not be equal. It depends on an atomic property called electronegativity. Electronegativity is the tendency for an atom to attract electrons within a bond. Greater attraction means greater electronegativity. Electronegativity increases as you move from left to right across the Periodic Table and decreases as you go down a group. Non-metals have a higher electronegativity than metals because their atoms have smaller radii – that is, the distance of the outer electrons from the nucleus is shorter. Therefore these smaller atoms have a greater pull on electrons because they are closer to the oppositely charged protons located in the nucleus. Values for electronegativity are located in Periodic Table below. Elements with great differences in electronegativity will form ionic bonds (no sharing); those with small differences will form covalent bonds, those in-between will form covalent bonds with ionic character or polar bonds (unequal sharing). Polar bonded compounds will show characteristics of both ionic and covalent compounds.



Occasionally, atoms will combine covalently but acquire additional electrons from another element that is not combining, or donate electrons to another atom. The result is a polyatomic ion. Polyatomic ions, although covalently bonded, are charged and can combine with other ions to form ionic compounds. Examples of polyatomic ions include: OH^- (hydroxide), CO_3^{-2} (carbonate), PO_4^{-3} (phosphate). Writing formulas for compounds with polyatomic ion, the ion must be put in put in parentheses before the subscript is added:

Ca^{2+} + $2NO_3^-$ ---- > $Ca(NO_3)_2$ correct

$CaNO_{32}$ incorrect

Common Polyatomic Ions

NH_4^+	ammonium	CO_3^{2-}	carbonate
H_3O^+	hydronium	HCO_3^-	hydrogen carbonate or bicarbonate
NO_2^-	nitrite		
NO_3^-	nitrate	ClO^-	hypochlorite
SO_3^{2-}	sulfite	ClO_2^-	chlorite
SO_4^{2-}	sulfate	ClO_3^-	chlorate
$S_2O_3^{2-}$	thiosulfate	ClO_4^-	perchlorate
HSO_4^-	hydrogen sulfate or bisulfate	$C_2H_3O_2^-$	acetate (or CH_3COO^- or $CH_3CO_2^-$)
OH^-	hydroxide		
CN^-	cyanide	MnO_4^-	permanganate
PO_4^{3-}	phosphate	CrO_4^{2-}	chromate
HPO_4^{2-}	hydrogen phosphate	$Cr_2O_7^{2-}$	dichromate
$H_2PO_4^-$	dihydrogen phosphate	O_2^{2-}	peroxide

Behavior of Ionic and Covalent Compounds
When They are Added to Water

It is common experience to dissolve the ionic compound NaCl (table salt) in water. The salt crystal, which is made up of an array of interlocking positive sodium and negative chloride ions, seems to disappear. What really happens is that the solvent water, which itself is a polar covalent molecule, pulls the ions apart and surrounds them so they cannot re-form the solid crystal. The individual Na^+’ ions and $C1^-$‘ ions are now dispersed in the water and are too small to be seen, although we can taste their presence. This process will be described in further detail in the Chapter on Solutions. Any soluble ionic compound will break-up into ions when placed in water. This is called dissociation.

Many polar covalent compounds, sugar for example, will also dissolve in water, but when they do, no ions are formed. The crystal simply breaks up into separate molecules that are dispersed throughout the water.

Most acids are polar covalent and undergo a reaction with water that produces ions. This production of ions by the action of water on acids is called ionization.

Non-polar covalent compounds, (compounds that do not have relatively positive and negative ends due to unequal sharing of electrons) generally do <u>not</u> dissolve in water. For example, gasoline which is a mixture of several non-polar covalent compounds, is <u>insoluble</u> in water.

C. <u>Misconceptions</u>

1. Chemical Bonds are actual physical connections.

   It is simple to explain a bond by showing sticks between two or more balls but this may also cause students to believe that a bond is an actual physical link. A chemical bond forms either from electrostatic attraction (ionic) which is a force or because electrons are shared (covalent) in which case the atoms have to stay close to one another to maintain the overlap of electron clouds. Since electrons are in constant motion they cannot form a firm connection. A covalent bond might be described better by thinking of two children who both want to play with a ball. They share by passing it back and forth. If either leaves or moves too far away, the sharing is impossible (the bond is broken). Neither one is connected to the other but they have a mutual interest, the ball (or electrons). This is another illustration of the limitation of models of physical phenomenon. It is a convenient way of having students, visualize how atoms bond together, but they can "create" wrong ideas!

2. We commonly teach that we eat food to "get" energy, and, that carbohydrates, proteins, and fats are big molecules that get "broken down" and release the energy that is stored in their bonds. But bond breaking <u>requires</u> energy. The energy that is obtained from food molecules comes from the <u>making</u> of new bonds – bonds that keep the smaller molecules that are the products of digestion together. Yes, food molecules are broken down, but that requires energy. It is the energy released when <u>new</u> bonds are made (which overall is greater than the amount required to break bonds) which result in the net gain in energy from metabolism.

3. "Metals like to lose electrons." In general, during chemical reactions metals which have fewer than four electrons in their outer shell lose these electrons and become positively charged. But these electrons just don't fall off the atom spontaneously. The negative e^- is attracted to the positive nucleus, and energy must be supplied to separate it. (The term for this energy is <u>ionization</u> <u>energy</u>). We also use words such as "like" when speaking of atoms as if atoms had some sort of will. The attributing of human qualities to inanimate objects is called ANTHROPOMORPHISM and may mislead students into thinking that atoms decide what to do.

D. <u>Warm-Up Exercises</u>

Before Lesson or Lab

1. Introductory Questions:

    1) How many different kinds of matter do you see around you?
    2) Are they all elements? If not, what could they be?
    3) How does the mater stay together?
    4) Why do atoms combine to form compounds?
    5) In what other ways do we use the word "bond?"
    6) Do you think it would take energy or release energy to break a bond? If energy is required to break a bond, what do you think happens to energy when you make a bond?

After Lesson or Lab

1. Show the two major kinds of bonding (ionic and covalent) by pretending students are atoms and are sharing or exchanging electrons (ball or any other item).

2. Make a concept may using <u>some</u> of the following terms: bond, electrons, compound, molecule, ionic bond, covalent bond, atom, ionic compound, covalent compound, ion, subscript, electronegativity.

3. Define vocabulary terms and give examples. Give definitions in words different from the text.

4. Use the position of elements on the Periodic Table to predict whether a compound is ionic or covalent from its formula.

5. Draw Lewis dot diagrams of some metals and non-metals. Using these diagrams – describe what would happen during bond formation between these atoms.

6. Imagine that you called "see" the individual particles in a glass of salt water or sugar water. What particles are present?

7. What would be the formula if atoms of the following elements combined?

    (HINT: a---- >c are ionic, d is covalent)

    1) Aluminum and fluorine
    2) Potassium and oxygen
    3) Calcium and iodine
    4) Oxygen and fluorine

E. Glossary

Anion	negatively charged ion formed by an atom when it gains electrons
Bond	a linkage that keeps atoms together
Cation	positively charged ion formed by the loss of electrons
Coefficient	number written in front of a formula which tells the number of molecules or formula units
Covalent bond	linkage between two atoms formed by sharing electrons
Electronegativity	measure of the tendency of an atom to attract electrons in a bond
Formula	expression showing the relative number of atoms of each element in a substance; it consists of symbols for the elements and subscripts; a symbol without a subscript means one atom
Formula unit	smallest piece of an ionic compound that contains the correct relative number of ions for each element in the compound
Group number	number assigned to a group or column on the Periodic Table, for the “A” groups it tells the number of valence or outermost electrons
Ion	an atom or group of atoms that has acquired a charge by gaining or losing electrons
Ionic bond	electrostatic force which holds ions together in an ionic compound; the attraction between oppositely charged ions
Ionic compound	compound composed of anions and cations held together by electrostatic force; usually a metal or non-metal
Law of Constant Composition	relation stating that the relative masses of elements in a compound is fixed; also known as the Law of Definite Proportions
Law of Multiple Proportions	relation stating that when two element, A and B, form two different compounds, the relative amounts of B which combine with A will vary by a ratio of small whole numbers
Lewis Dot Structure	representation of atoms that include the element’s symbol and dots for the valence electrons
Molecule	smallest unit of a covalent compound with all chemical properties of the compound

Octet rule	principle that states that atoms tend to have eight electrons on their outermost shells
Polar bond	a covalent chemical bond where electrons are unequally shared; molecule has slightly positive and negative ends
Polyatomic	two or more atoms covalently bonded that have a positive or negative charge
Subscript	number in a formula indicating the number of atoms of each element in a compound
Valence electrons	electrons in the outermost shell of an atom

CHAPTER 4

Chemical Reactions

B. Background

A chemical reaction (or chemical change) is a rearrangement of atoms in which chemical bonds are broken, or made, or both. It produces changes in the chemical and physical properties of the substances involved. We symbolize the reaction by using a chemical equation. In a chemical equation:

reactants(s) --------> product(s)

The arrow represents the phrase "react to form."

As the word 'equation' implies, 'things' must be equal on both sides of the arrow. Things here refers to atoms. Since the Law of Conservation of Matter states that matter cannot be created nor destroyed, all atoms must be accounted for. When a reaction is written as an equation, it may already conform to the Law:

Chemical equation: $CaCO_3$ --- > CaO + CO_2
Word equation: calcium carbonate reacts to form calcium oxide and carbon dioxide
1 Ca, 1 C, 3O = 1 Ca, 1 C, 3 O

In some equations, the number of atoms is not equal. For example:

Chemical equation: H_2O --- > $H_2 + O_2$
Word equation: water reacts to form hydrogen and oxygen

Note: Both hydrogen and oxygen must be written with a subscript of 2 because that is the way they exist in nature – as diatomic molecules not as single atoms.

Equations that are not equal must be balanced by adjusting the proportions of each reactant and/or product:

$2H_2 + O_2$ --- > $2H_2O$
4 H, 2 O = 4 H, 2 O

The new equalized statement is called a balanced chemical equation. The number added in front of each compound is called a coefficient.

C_3H_8 + O_2 ---> CO_2 + H_2O unbalanced

C_3H_8 + $5O_2$ ---> $3CO_2$ + $4H_2O$ balanced

3 C, 8 H, 10 O = 3 C, 8 H, 10 O

You can NEVER change a subscript to balance an equation because changing a subscript changes the identity of the compound. You then change the nature of the reaction you are trying to symbolize.

The coefficients express the proportions in which reactants and products are consumed and produced. For:

$$C_3H_8 + 5O_2 \text{ --- > } 3CO_2 + 4H_2O$$

1 molecule of C_3H_8 reacts with 5 molecules of O_2 produces 3 molecules of carbon dioxide and 4 molecules of H_2O. Note that you start with 6 molecules and produce 7, but that is fine as long as the number of atoms is the same. Keep in mind that atoms are reassembling into different combinations.

Since individual atoms or molecules are too small to handle, the coefficients can also indicate moles. So this same equation can be read as: 1 mole of C_3H_8 reacts with 5 moles of O_2 to yield 3 moles of CO_2 and 4 moles of H_2O.

Often, an equation may contain extra symbols that give additional information about the reaction.

$$C_3H_8\,(g) + 5O_2\,(g) \text{ --- > } 3CO_2\,(g) + 4H_2O(1) \;+ \Delta$$

In the above reaction, (g) represents gas and indicates that C_eH_8, O_2 and CO_2 are all gases, while H_2O is a liquid (1).

g or ↑ = gas	aq = aqueous (dissolved in water)
1 = liquid	Δ = heat
s or ↓ = solid	

Occasional symbols are used above or below the arrow to indicate what was done to foster the chemical change.

Δ = heat hν = light

Types of Chemical Reactions

There are a number of general types of chemical reactions which have corresponding general types of equations to express them.

1. In a <u>combination</u> or <u>synthesis</u> <u>reaction</u>, two substances combine to form one. The equation to describe this is also called combination.

   A + X --- > AX

$C + O_2$ --- > CO_2	Two elements can combine to form a compound
NH_3 +HC1 --- > NH_4Cl	Two compounds can combine to form a larger compound

2. In <u>decomposition</u> or <u>analysis</u> <u>reactions</u>, one reactant is broken down into two or more substances. The type of equation for this is also called decomposition or analysis.

$$AX \text{ --- > } A + X$$

$CaCO_3(s)$ --- > $CaO(s) + CO_2\,(g)$	A compound breaks down into 2 smaller compounds
$2H_2O(1)$ --- >$2H_2\,(g) + O_2\,(g)$	A compound breaks into elements

3. An <u>oxidation-reduction</u>, or <u>redox</u>, <u>reaction</u> is one where electrons are transferred from one substance to another. You can recognize that this type of reaction has taken place if the charge of an element changes from one side of an equation to the other. One way to represent redox reactions is by using <u>single</u> <u>replacement</u> <u>equations</u>. Single replacement reactions are specific redox reactions that involve an element and a compound as reactants and a different element and different compound are the products.

$$A + BX \text{ --- > } AX + B$$

$Cu\,(s) + 2AgNO_3\,(aq)$ --- > $Cu(NO_3)_2\,(aq) + Ag\,(s)$

copper has no charge --- > copper has + 2 charge,
silver has a charge of +1 --->. silver has no charge

$Mg\,(s) + 2HC1(aq)$ --- > $MgCl_2(aq) + H_2\,(g)$

Mg has no charge, H is +1 --- > Mg has +2 charge,
Hydrogen has no charge

4. A <u>double replacement</u> reaction involves two ionic compounds that exchange ions. It is as if the substances "swap partners".

$$NaCl\,(aq) \quad + \quad AgNO_3\,(aq) \quad \text{---->} \quad AgCl\,(s) \quad + \quad NaNO_3\,(aq)$$

Many reactions occur in aqueous medium (in water) with one or more ionic compounds. Remember that ionic compounds dissociate or separate into their ions when they dissolve in water. It is often convenient to write ionic equations (i.e., equations that show the ions present) to more accurately indicate what is occurring.

Equation:

$Cu(s) + 2AgNO_3\,(aq) \text{ --- > } Cu(NO_3)_2\,(aq) + 2Ag\,(s)$

Ionic Equation for the above reaction:

$Cu(s) + 2Ag^+\,(aq) + 2NO_3^-\,(aq) \text{ --- > } Cu^{2+}\,(aq) + 2NO_3^-\,(aq) + 2Ag\,(s)$

If we eliminate any item that is the same on both sides we get a net ionic equation.

Net Ionic Equation:

$Cu(s) + 2Ag^+\,(aq) \text{ --- > } Cu^{2+}\,(aq) + 2Ag\,(s)$

Net ionic equations are useful for looking at a type of double replacement reaction called a precipitation reaction. In a precipitation reaction, ions combine to form an insoluble product, a precipitate.

AX + BY --- > AY + BX (AY or BX is insoluble in water)

Two compounds react to form 2 different compounds.

$AgNO_3\,(aq) + NaBr\,(aq) \text{ --- > } AgBr\,(s) + NaNO_3\,(aq)$

Ionic Equation:

$Ag^+ + NO_3^-\,(aq) + Na^+\,(aq) + Br^-\,(aq) \text{ --- > } AgBr\,(s) + Na^+\,(aq) + NO_3^-\,(aq)$

Net Ionic Equation:

$Ag^+(aq) + C1^-(aq) \text{ --- > } AgC1(s)$

When looking at a double replacement equation, you can determine which of the products is a precipitate by referring to the solubility rules, see the table below.

Soluble Compound	**Exceptions**
Almost all salts of sodium (Na^+), potassium (K^+) and ammonium ions (NH_4^+)	
All salts of chloride (Cl^-), bromide (Br^-) and iodide (I^-)	Halides of silver (Ag^+), mercury (I) (Hg_2^{2+}) and lead (II) (Pb^{2+})
Salts of nitrates (NO_3^-), chlorates (ClO_3^-), perchlorates (ClO_4^-), and acetates ($C_2H_3O_2^-$)	
Salts of fluoride (F^-)	Fluorides of Group IIA metals (Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}) and lead (II) (Pb^{2+})
Salts of sulfate (SO_4^{2-})	Sulfates of barium (Ba^{2+}), strontium (Sr^{2+}) and lead (II) (Pb^{2+})
Insoluble compounds	**Exceptions**
Salts of carbonate (CO_3^{2-}), phosphate (PO_4^{3-}), oxalate ($C_2O_4^{2-}$), chromate (CrO_4^{2-}), sulfide (S^{2-}), hydroxide (OH^-), and oxide (O^{2-})	Salts of ammonium ion (NH_4^+) and Group IA metals (Li^+, Na^+, K^+, Rb^+)

Another type of double replacement reaction is called neutralization reaction.. In neutralization, an acid (H^+ (acting as a non-metal) or H_3O^+ (negative polyatomic ion)) and base (containing hydroxide (OH^-)) come together to form an ionic compound and water.

$$AX + BY \text{ ---> } AY + BX$$

$$HCl\,(aq) + KOH\,(aq) \text{ ---> } KCl\,(aq) + H_2O\,(1)$$

Ionic Equation:

$$H^+\,(aq) + C1^-\,(aq) + K^+\,(aq) + OH^-\,(aq) \text{ ---> } K^+\,(aq) + C1^-\,(aq) + H_2O\,(1)$$

Net Ionic Equation:

$$H^+\,(aq) + OH^-\,(aq) \text{ ---> } H_2O\,(l)$$

Reversible Reactions

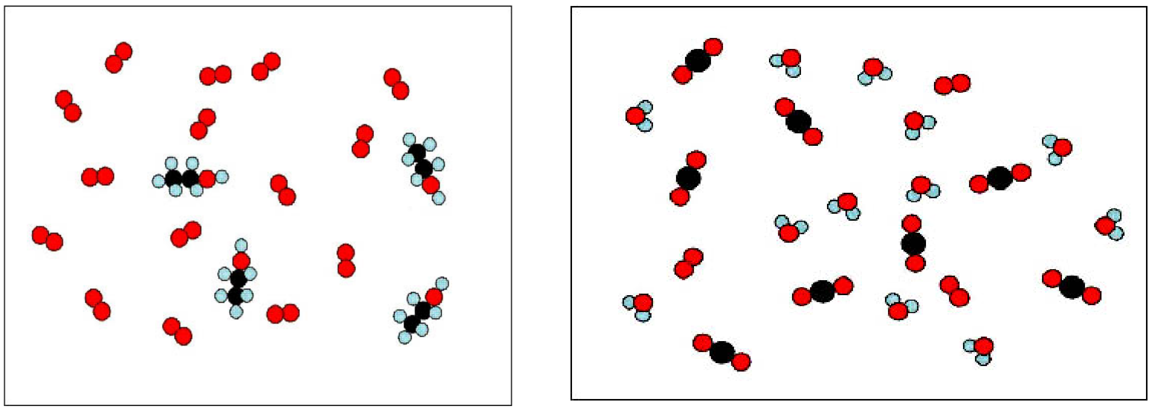
No chemical reaction goes all the way to completion. However, most come close with amount of product far exceeding amount of unused reactant. We generally ignore the unused reactant and say the reaction is complete. Some reactions are only partially complete and in fact may easily reverse depending on the conditions. These reactions are shown with a double arrow:

$$CaCO_3\,(s) \rightleftharpoons CaO\,(s) + CO_2\,(g)$$

$$H_2\,(g) + I_2\,(g) \rightleftharpoons 2\,HI\,(g)$$

At any time, you will find products and reactants in the reaction container. The system containing the reactants and products eventually establishes a dynamic equilibrium. The rate of the forward and reverse reactions, are equal at equilibrium. If you stress the system (by adding or removing product or reactant or changing the temperature), the equilibrium is disturbed but will eventually be reestablished. Le Chatelier's Principle states that when a stress is applied to a system in equilibrium, the system will react in such a way as to overcome the stress. Therefore, if you add the stress of extra reactant, the reaction system adjusts in order to remove this extra reactant causing more of the product to be formed.

Understanding Chemical Changes Using Particles in a Box



BEFORE **AFTER**

We write reactions using symbols because it is convenient, but if we want students to have a better understanding of what is happening in a chemical change it is better to use "particles in a box" graphics. In the boxes above you can see what particles exist before and after the change. The reaction that is taking place is:

$$C_2H_6O \quad + \quad 3\ O_2 \quad \text{---->} \quad 2\ CO_2 \quad + \quad 3\ H_2O$$

You may notice that there are some molecules in the AFTER box that were in the BEFORE box. These are molecules of the reactant that were left over. We DO NOT write these as products in the chemical equation but just acknowledge that both reactants may not be consumed in a change.

C. Misconceptions

1. If you cannot see anything happening when two reactants are mixed, no reaction is taking place.

   Some reactions, particularly neutralization, have reactants and products which are colorless and water soluble. We do not see a change but it occurs nonetheless. If you mix solutions of sodium hydroxide (lye) and hydrochloric acid (stomach acid), water and table salt are the products. But the salt dissolves in the water and so the overall appearance has not changed. You can however "feel" the heat generated by touching the reaction container.

2. When attempting to describe a liquid mixture many students will say that it is “white” – when it is actually colorless and clear. The term “clear” means transparent. Consequently, solutions such as copper sulfate – can have a color (blue) and be clear at the same time. If a liquid is not clear, the proper term is “cloudy,” or if a solid settles to the bottom, a “precipitate” is said to have formed.

3. When you write a chemical equation both the reactants and products are shown so some students will assume that all these materials are present at the end. Although reactions rarely go to completion it is still important for students to understand that one or both of the reactants will be gone when the reaction finishes unless it is a reversible reaction.

D. Warm-Up Exercises

Before Lesson or Lab

1. If elements combine to form compounds, can compounds combine or react with compounds? How?

2. Tara makes a solution of baking soda and mixes it with vinegar. The mixture foams up immediately. What happening? Are the vinegar and baking soda changed? How could you tell?

After Lesson or Lab

1. Predict what will happen for each reaction (use solubility rules).

$NaC1\ (aq) + AgNO_3\ (aq) \dashrightarrow$
$HC1\ (aq) + Mg(OH)_2\ (aq) \dashrightarrow$
$HgC1_2\ (aq) + NaI\ (aq) \dashrightarrow$

2. a) Tim had 3 solutions from a 6-bottle experiment that were all clear and colorless and the labels fell off. He knows they must be $Pb(CH_3COO)_2$, KI, and $HgC1_2$. In a flash he sees a way to get answer. Do you?

   b) Tim labels the bottles 1, 2, and 3. Mixing 1 and 2 gives an orange solid; 1 and 3 given a bright yellow solid; 2 and 3 have no reaction. What are 1, 2 and 3?

   c) Do you recall the copper sulfate solution that was used in the experiment with the nail? How would you describe it? What is the difference between the terms colorless and clear? Give examples.

3. Construct a concept may with the terms: matter, chemical reaction, chemical equation, coefficient, reactant, product, atoms, Law of Conservation of Matter, balancing.

4. Give examples of different types of chemical reactions from everyday life. HINTS:

Silver tarnishing $2Ag + S \text{ ---> } Ag_2S$ Combination

Taking an antacid tablet

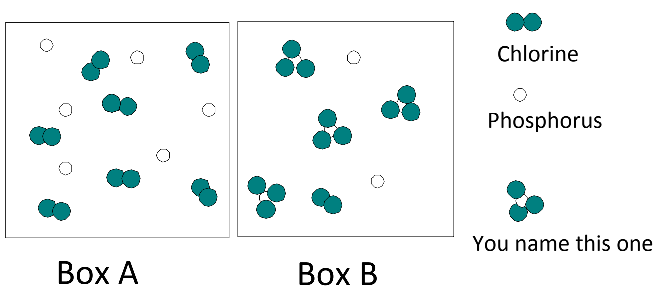
$$CaCO_3 + 2HCl \text{ ---> } CaCl_2 + \underset{\swarrow\quad\searrow}{H_2CO_3}$$ Double replacement

Tums stomach acid $H_2O + CO_2$ (burp)

Releasing bubbles from carbonated soda $H_2CO_3 \text{ ---> } H_2O + CO_2$ Decomposition

5. Define vocabulary terms. Use word different from those used in the textbook.

6. Write and balance the equation for the reaction shown in the boxes below.



E. Glossary

Aqueous — dissolved in water

Balanced equation — chemical equation in which reactants and products contain the same number of atoms of each type

Chemical equation — expression that symbolizes a chemical reaction (qualitatively and and quantitatively – what is there and how much of it is there)

Chemical reaction — a rearrangement of atoms that produces changes in physical and chemical properties of substances involved; chemical change in which new materials are formed.

Coefficient — large number preceding chemical formulas in a chemical equation

Combination reaction — reaction where two elements or an element and compound or two

	compounds combine to form one product
Decomposition reaction	reaction where one reactant forms two or more products
Double-replacement equation	equation that has the general form: AX + BY ---> AY + BX
Dynamic equilibrium	state where two opposing processes or reactions are occurring at the same time and same rate
Net-ionic equation	chemical equation obtained by omitting unchanged ions (spectator ions) from an ionic equation
Neutralization	reaction of an acid and base to form a salt and water
Oxidation-reduction	reaction that involves transfer of electrons from one substance to another; also called redox
Precipitate	solid that forms when two solutions are mixed, usually as a result of a chemical reaction
Precipitation reaction	reaction in which a precipitate is a product
Product	what is formed as a result of a chemical reaction
Reactants	substances are present at the beginning of a chemical reaction and subsequently undergo a change in identity
Redox	see oxidation reduction
Single replacement reaction	chemical equation that has the general form A + BX ---> AX + B

F. Additional Resources

Background on balancing equations
http://www.visionlearning.com/library/module_viewer.php?mid=56

A tutorial to assist in balancing equations
http://antoine.frostburg.edu/chem/senese/101/kits/kit_chemical_equation.html

An activity to help learn how to balance equations
http://www.middleschoolscience.com/balance.html

CHAPTER 6

Physical States of Matter

B. Background

The three physical states of matter we are familiar with are: gas, liquid, and solid. We know that water, H_2O, exists as a solid, ice, below O^0C, as a liquid at room temperature, and as vapor or in the gaseous state (causing various degrees of humidity). In all three states, the identity of the compound water stays the same; therefore, the changing of state is not a chemical change. It is a physical change. Each physical state has different, unique, characteristic properties, but the chemical properties of a substance remain the same no matter what the state.

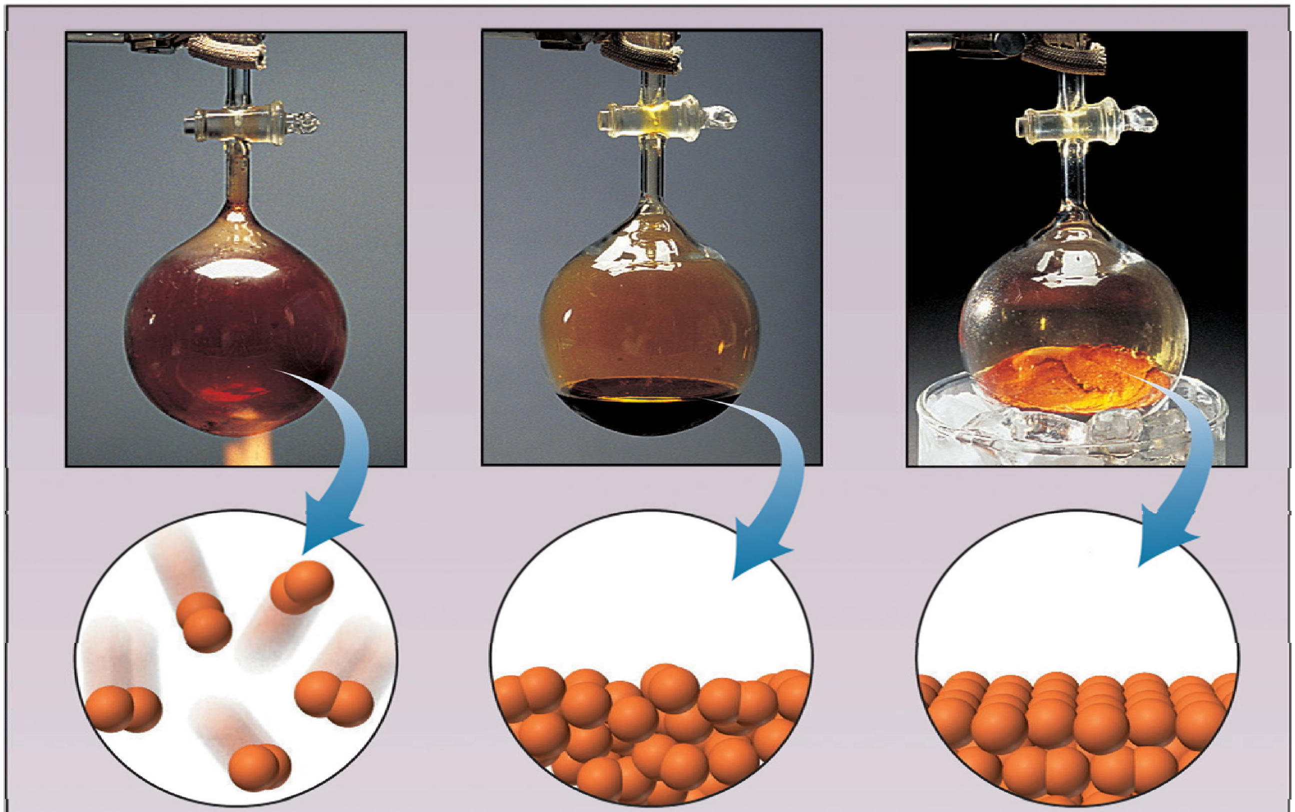
Solids generally maintain their shape and volume no matter what their location.

Liquids assume the shape of their containers. However, like solids, they maintain a fairly constant volume.

Gases do not maintain a definite shape or a definite volume. They can expand or contract to assume the shape and volume of the container. They completely fill their containers.

Some Common Characteristics of Gases, Liquids, and Solids

GASES	LIQUIDS	SOLIDS
1. No definite shape (fill containers completely)	1. No definite shape (assume shape of container)	1. Definite shape (resists deformation)
2. Compressible (can be squeezed)	2. Incompressible	2. Incompressible
3. Low density	3. Intermediate –high density	3. Intermediate-high density
4. Fluid	4. Fluid	4. Not fluid
5. Diffuse rapidly through other gases	5. Diffuse through other liquids	5. Diffuse very slowly through solids
6. Extremely disordered particles; much empty space; rapid, random motion in three directions	6. Disordered clusters of particles; quite close together; random motion in three dimensions	6. Ordered arrangement of particles; vibrational motion only; particles very close together



A Gas: Molecules are far apart and fill the available space

B Liquid: Molecules are close together but move relative to each other

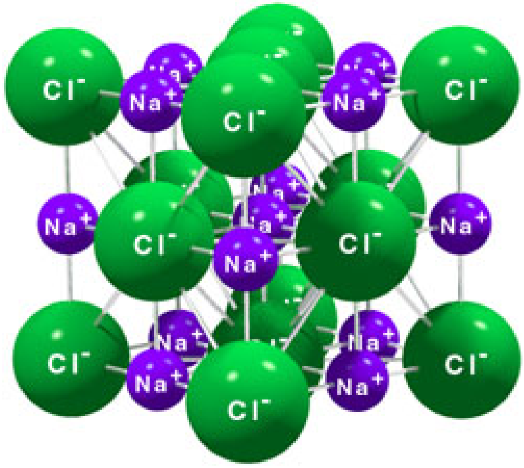
C Solid: Molecules are tightly packed in a regular array and move very little relative to each other

While the three physical states of matter have their unique characteristics, matter can exist in each state and can be converted from one physical state to another through familiar physical changes, such as melting, condensation, and evaporation.

Since matter can exist in all three states, some pertinent questions are: What is holding particles of a given substance together in a rigid crystal lattice in their solid state? What keeps them together in the fluid liquid state? What causes particles to be apart in the gaseous state? What is needed to convert matter from one state to another?

What is the Nature of Solids?

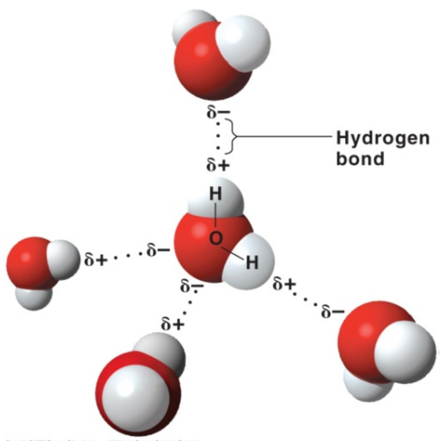
The rigidity of solids is due to the fact that their unit pieces, molecules in the case of covalent compounds, and ions in the case of ionic compounds, are held together by intermolecular forces (IMFs) – forces between molecules- thereby keeping the molecules or ions firmly in place. There is some motion among the solid particles; they actually are vibrating (the speed of which increases as temperature is increased) but there isn't any translational motion. They



http://www.chemistry.wustl.edu/~edudev/LabTutorials/Water/PublicWaterSupply/images/nacl.jpg

do not move around in relation to each other. In <u>ionic compounds</u>, it is the attraction between the oppositely charged ions (<u>electrostatic force</u>) that is keeping the solid together. The term intermolecular forces, is probably not appropriate to use with ionic solids, as there really aren't any molecules present. A simplified model of a salt crystal would show an array of alternating Na^+'s and Cl^- 's in a cubic shape. There are <u>no</u> individual NaCl molecules as all the ions are interconnected. We instead refer to the smallest piece of an ionic compound as a <u>formula unit</u>.

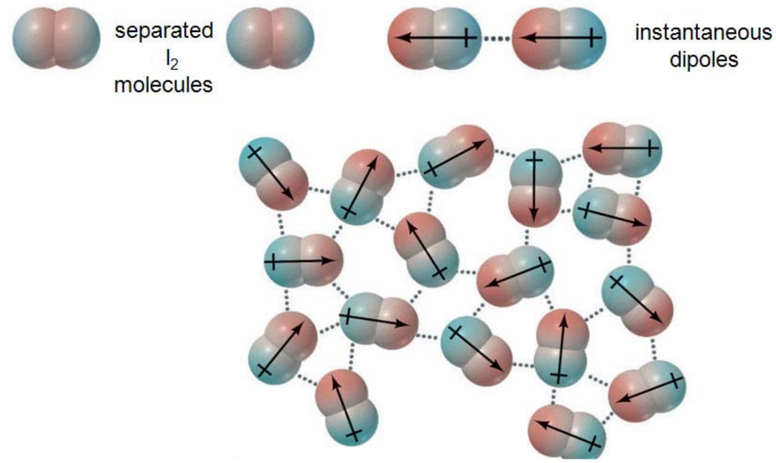
<u>Covalent compounds</u>, on the other hand, do exist as individual molecules. Water, for example, is polar covalent and has relatively positive and negative ends referred to as <u>poles</u>. These polar molecules called dipoles are like miniature magnets and the positive end of one water molecule is attracted to the negative end of another. This attraction between the oppositely charged ends of polar molecules is a type of intermolecular force called <u>dipole-dipole interaction</u> and is one of the forces in operation that holds ice molecules together in the solid state.



https://eapbiofield.wikispaces.com/file/view/03_02BondWater-L.jpg

In water, the tiny hydrogen atoms are directly bonded to the much larger (electronegative) oxygen atom. The oxygen has a very strong pull on the e^-'s that are being shared making the hydrogen more positive than is usual in polar bonds. Consequently, there is another intermolecular force in operation in ice, called <u>HYDROGEN BONDING</u>, in which the hydrogens from one water molecule are attracted to the oxygen atom of neighboring atoms. This attraction is so stabilizing (energy releasing) that water molecules arrange themselves in ice crystals to maximize the number of these attractions that can form. This phenomenon is responsible for the fact that ice, contrary to the usual situation with solids, is less dense than liquid water. Hydrogen bonding is also at work in the molecule DNA. In DNA some hydrogen atoms are attached directly to nitrogens, and the molecule arranges itself so that as many hydrogen bonds can occur between H's and N's as possible resulting in the spiral-like (helical) shape of the famous molecule. Hydrogen bonding occurs between molecules when hydrogen is attached directly to oxygen, nitrogen or fluorine.

There are some solids that are made up of non-polar covalent molecules. Solid iodine is an example. Iodine is diatomic, I - I, where two iodine atoms share a pair of electrons and form the iodine molecule. It is totally non-polar, each atom has an equal pull on the electron pair. So there are no dipole-dipole interactions. What then is keeping the iodine molecules together in solid iodine? It is helpful to remember and imagine that electrons form clouds of negative charge around nuclei, so 2 iodine molecules can be pictured as they are in the diagram below. When one I_2 molecular approaches, the electron clouds repel each other and become distorted exposing, for an instant, more of the nuclei. This allows the e^-'s from one molecule to become attracted to the nuclei of another molecular. This is a weak temporary intermolecular force called London force or dispersion forces, and it is what keeps iodine in the solid state. It is also the intermolecular force at work in dry ice, solid CO_2, which is also a non-polar covalent molecular.



What is Happening When Solids Change to Liquids, or Liquids Change to Solids?

We all know that leaving ice out in the sun will cause it to change to liquid water. Heat is a form of energy and this energy is used to overcome the intermolecular forces that are holding the solid together. The change from the solid to the liquid state is called melting, and the particular temperature at which it occurs is the melting point. Since the strength of intermolecular forces varies from solid to solid, melting points also vary. (H_2O melts at 0^oC). What can you know about the relative strength of IMF's of 2 solids, if one has a higher melting point than the other? The solid with the higher melting point has the stronger IMF's. Since melting requires heat it is referred to as an endothermic process.

The process of changing from a liquid to a solid, called freezing, involves the release of energy (IMF's are like bonding between molecules – and remember bond making is energy releasing) and so it is an exothermic process. The freezing point is the same as the melting point only the direction of the energy transfer is different. (H_2O freezes at 0^oC).

When the rigid structure of the solid is broken, particles are freed to move around in relation to one another which is what is happening in the liquid state. Forces of attraction still keep the particles near each other, but they are not locked in place thus explaining why liquids are fluid, or are capable of flowing from one place to another.

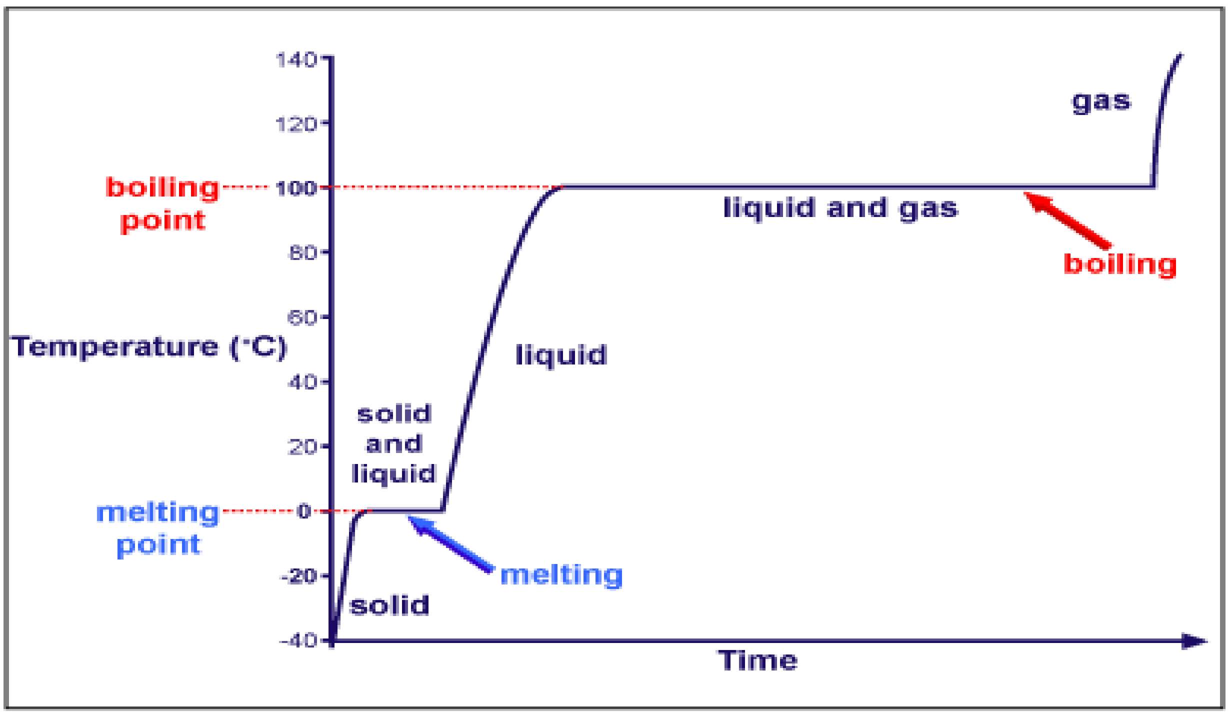
What Happens When a Liquid Changes to a Gas?

Once the liquid state has been achieved, and additional heat energy is added, individual particles gain sufficient energy to break away – and enter the gaseous state. This process is known as evaporation, and occurs at the surface of liquids. When the temperature reaches what is known as the boiling point, evaporation occurs very rapidly and extends to particles below the surface. So those bubbles that appear throughout a liquid during boiling are actually gas of whatever liquid is being boiled. Gas particles are independent of one another, no intermolecular forces exist between them, and they are very far apart, and moving in random straight-line directions.

In order to change back to a liquid, a process known as condensation, the gas molecules have to come closer together and re-make attachments between them. This is energy releasing (exothermic). When steam condenses on your hand, the burning sensation is caused by the release of all that energy onto your skin. The condensation point or condensation temperature is the same as the boiling point.

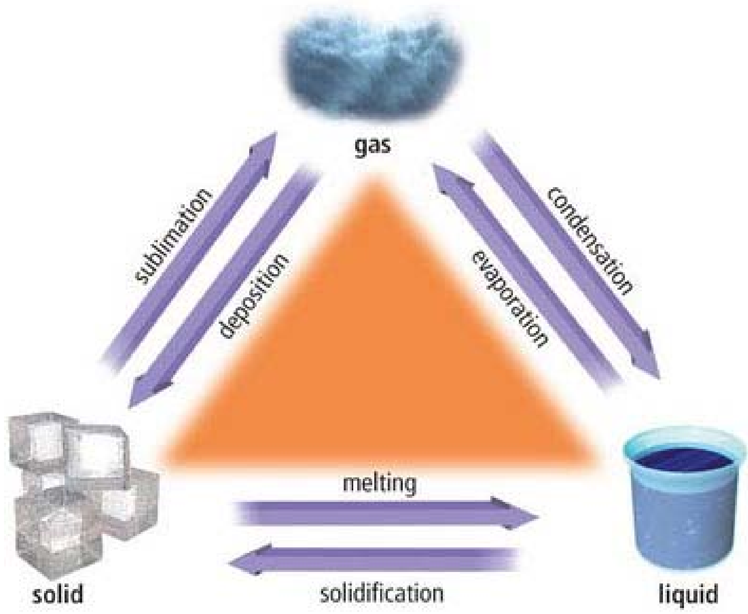
Note: Certain solids that have very weak intermolecular forces change directly from the solid state to the gaseous state, a process known as SUBLIMATION. Iodine and dry ice are examples of such solids. Remember it is only the weak London forces that hold these solids together.

The diagram below shows the changes in the physical states of water with the increase in temperature. This type of diagram is called a heating curve. If you gradually cool a material as it changes phases you generate a cooling curve.



http://www.bbc.co.uk/schools/ks3bitesize/science/images/sci_dia_21.gif

Through all these changes the water, H_2O, molecules remain intact; the only change is the distance between them and the subsequent physical characteristics. In the case of water, sugar, alcohol and other similar compounds, the particles in the crystal lattice and in the liquid are molecules. In the case of table salt (NaCl), lye (NaOH), and other salts, bases and some acids, the particles occupying the points in the crystal lattice and moving in the liquid are ions. Gases of these compounds would also consist of individual ions.



Common Examples of Phase Changes

EVAPORATION (Liquid --- > Gas)

- Feeling cool after getting out of a pool on a windy day
- Water level dropping in the fish tank at home
- The tea kettle whistling (rapid evaporation throughout the liquid)

http://images.inmagine.com/img/ojoimages/oj060/pe0060470.jpg

CONDENSATION (Gas --- > Liquid)

- Dew forming on a car windshield
- Fogging the mirror after a hot shower
- Seeing your breath on a cold morning

FREEZING (Liquid --- > Solid)

- Ice forming
- Lava from a volcano turning to rock

MELTING (Solid --- > Liquid)

- Snow disappearing on a warm day
- Heating butter or margarine to fry an egg

CONDENSATION (DEPOSITION) (Gas --- > Solid)

- Frost forming on a car windshield

SUBLIMATION (Solid --- > Gas)

- Dry ice disappearing
- Frozen clothes drying outside on winter day

C. **Misconceptions**

1. Boiling water continues to increase in temperature as heat is added.

   It is not difficult to understand why someone would think that temperature goes up continuously as heat is being applied. But the heat that is being added to water that is already boiling is being used to overcome the intermolecular forces of the liquid and separate molecules into the gaseous state. The amount of heat necessary to change 1 gram of water to 1 gram of steam is 540 calories (heat of vaporization), while it only takes 1 calorie to raise 1 gram of liquid water 1^o centigrade (specific heat).

   It is also true that a glass of water with ice in it will remain at the freezing point of H_2O, 0^oC, until all the ice is melted, even if it is placed in the sun. Here again, heat energy that is added is used to change the solid to liquid, not to raise the temperature. It takes 80 calories to change 1 gram of ice at 0^oC to 1 gram of water at 0^oC (heat of fusion).

2. When students are asked to represent the three states of matter using dots to represent particles, some of their misconceptions come to light. A common error is that they leave huge spaces between liquid particles, and fail to put the crystalline solid particles in any sort of arrangement. They will also show gas particles close together or all clustered either at the top or bottom of the container instead of dispersed throughout the container. Some will draw squiggly lines for gases and liquids that betrays their lack of "commitment" to the idea that these phases of matter are actually made of individual

3. "Water undergoes a chemical change when it is boiled." When students are asked to identify the "gas" coming from a boiling beaker of water, many of them say "hydrogen" and "oxygen." They evidently know that water consists of H2O, but erroneously interpret the apparently violent process of boiling as ripping the water molecules apart. Testing the vapors of steam using a glowing splint (it relights in oxygen) and a burning splint (it makes a popping sound in hydrogen) will disprove this view.

4. Some students identify the bubbles in boiling water as "air" rather than gaseous water (during boiling water undergoes rapid evaporation below the surface). An approach to counter this error is to ask students how all that air got into the water in the first place.

D. **Warm-Up Exercises**

1. What are some observations you can make about solid ice, liquid water and steam. If you have bionic eyes to see inside, what would you see? Draw pictures of each using little circles to represent the particles.

2. Explain the following terms: Give many examples of each. How can you prove that they are all physical changes?

   Evaporation
   Condensation

Boiling
Melting
Freezing

2. What is the difference between melted sugar and hot sugar/water solution? How are they the same?

3. Is there such a thing as gaseous iron? What might it look like?

4. Why is it correct to say that lava freezes when it comes out of a volcano?

5. Have each member of the class represent a particle – atoms, ions, or molecules. Start out by having them arrange themselves to represent a solid. Then add heat. They should vibrate more rapidly and then move from place to place as they "liquefy" – but remain shoulder to shoulder. Add more heat. As "gases" the particles will be moving and vibrating but will now be far apart.

Gases

Whenever we discuss gases, we have to keep in mind that gas molecules are separated by large amounts of empty space. Because of all the space between them, gases can be squeezed or compressed. For the same reason, since gaseous molecules are moving freely through space, gases can expand. Expansion and compression of gases change only the amount of empty space between the molecules. An "empty" room is not empty. It contains invisible molecules of gases, such as oxygen, O_2, carbon dioxide, CO_2, water, H_2O, and others. But the empty space between gas particles is true emptiness since there is "nothing" between them.

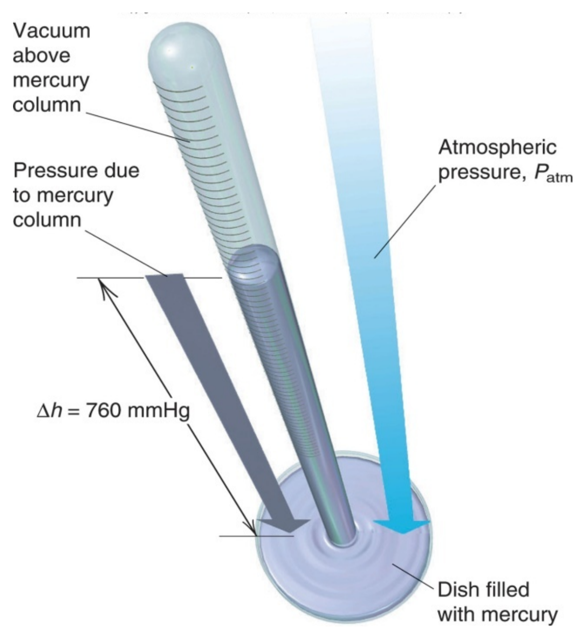
Particles of gases are moving through space in rapid, random, straight line paths. They possess kinetic energy (energy an object possesses as a result of its motion). As the temperature is raised the particles move faster, so temperature is directly proportional to the kinetic energy of the gas.

When describing the physical behavior of gases we must consider physical conditions such as volume, pressure, temperature and how many gas particles we have. Since these conditions can be altered, they are referred to as variables.

Volume, V, is the space occupied by a given gas. Units for volume are liters, L, milliliters, mL, and cubic centimeter, cm^3. Since a gas completely fills its container, the volume of a gas is equal to the volume of the container that holds it. Remember:

$$1\text{ L} = 1000\text{ mL} \qquad 1\text{ mL} = 1\text{cm}^3$$

Pressure, P, is force per unit area. As the gas molecules move freely and chaotically through the space in the container, they collide with the walls of the container thus exerting PRESSURE on them. Pressure is expressed in many different units. The ones most frequently used are atmosphere, atm; millimeters mercury, mmHg, and/or torr. The instrument used to measure atmospheric pressure is the barometer (invented by Torricelli). You will hear the "pressure" described in weather reports in units of inches. It is just the English system standard for length – and can be converted to mm Hg.



1 atm = 760 mmHg = 76 cmHg = 14.7 psi

1 mmHg = 1 torr

Temperature, T, measures degree of "hotness." The two temperature scales used when dealing with gases are degrees Celsius, ^oC, and the Kelvin scale, K. In calculations involving temperature, the Kelvin scale must be used, as 0 kelvins means no kinetic energy, whereas 0^oC is not a true zero value (since you can have negative temperature in this scale).

$$0^oC = 32^oF = 273\ K$$
$$100^oC = 212^oF = 373\ K$$
$$K = 273 + {}^oC$$

Number of Moles, n, refers to how many gas particles are present. One mole of a gas equals its molar mass in grams. If you had 2 grams of hydrogen which is one mole, it contains 6.02×10^{23} hydrogen molecules.

$$\#\ moles = \frac{grams}{molar\ mass}$$

The Gas Laws (Boyle's, Charles', Guy-Lussac's, Avogardro's Principle, Ideal Gas Law, Combined Law) are expressions of the relationships among the gas variables.

Gas Laws

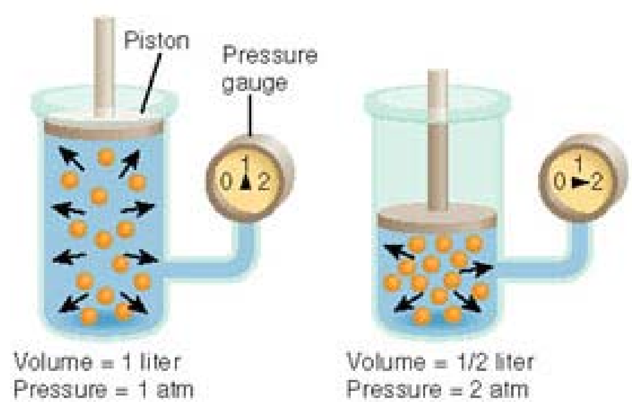
Boyles' Law – P/V Relationship

Consider an "empty" syringe or a syringe containing air. When you apply pressure (P is increased) the piston moves down and the volume decreases. When you release or decrease pressure the piston moves up and the volume increases. Pressure and volume are indirectly proportional, as one gets larger, the other gets smaller. This same relationship allows gases to be compressed in steel tanks for diving and resuscitation. This relationship between pressure, P, and volume, V, at a given temperature with a given amount of gas (T and n are constant) was investigated and mathematically interpreted by Robert Boyle in the 1600's. Consequently, it is referred to as Boyle's Law.

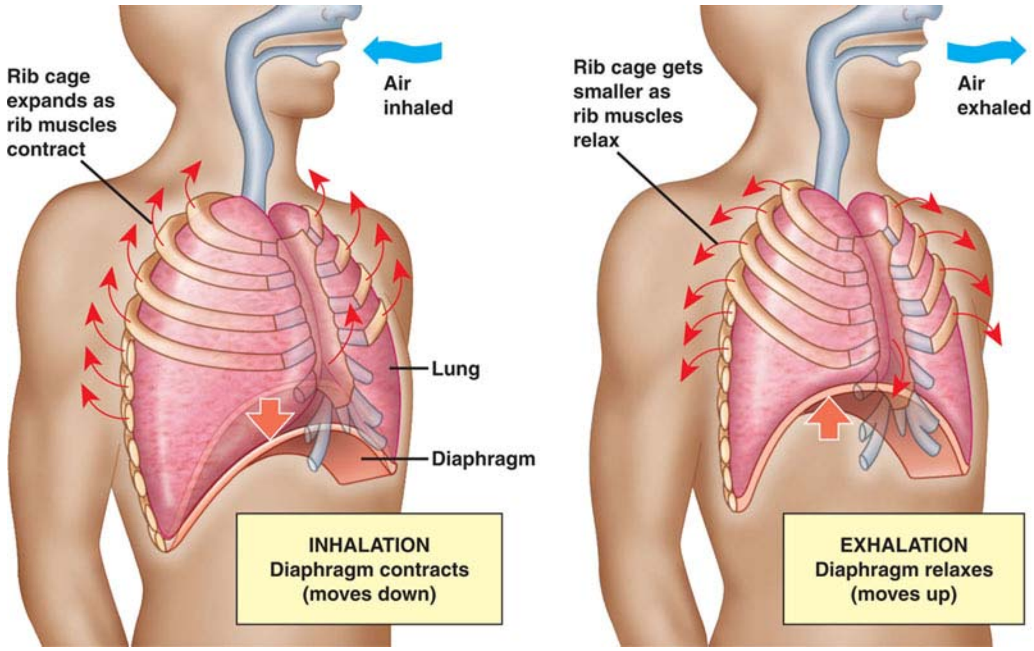
Mathematically expressed this relationship is Boyle's Law:

$$P = constant \times \frac{1}{V} \text{ or } PV = constant$$

$$P_1V_1 = P_2V_2 = P_3V_3 = \ldots = P_nV_n = constant$$



Breathing, the action of your lungs, is an example of Boyle's Law.



http://kvhs.nbed.nb.ca/gallant/biology/negative_pressure_breathing.html

air flows into lungs   air flows out of lungs

In the first figure above, the diaphragm is relaxed, V of lungs is increased, and pressure inside lungs falls below the pressure of the atmosphere. Since air will flow from an area of high pressure to an area of lower pressure, air rushes into the lungs. You inhale.

In the second figure, the diaphragm contracts decreasing volume of the lungs; the pressure goes up above atmosphere pressure. Air rushes out. You exhale.

Charles' Law – Temperature/Volume Relationship

We know that a balloon filled with gas exposed to heat (sun) will expand. The higher the temperature, the greater the volume, (provided that the outside pressure remains unchanged and that we deal with the same amount of gas of same number of gas molecules). If the same balloon is placed in the refrigerator, it will shrink. The volume of a gas is directly proportional to the Kelvin temperature, as one gets larger, the other also gets larger. <u>Charles'</u> <u>Law</u> is the mathematical expression that relates volume to temperature.

$V = \text{constant} \times T$ or $\frac{V}{T} = \text{constant}$

(Note that the temperature must be expressed in K.)

$\frac{T_2}{T_1} = \frac{V_2}{V_1}$ (P = constant, number of gas molecules = constant)

An example of the temperature/volume relationship is the rising of cake batter.

-------->
OVEN

initial batter with
trapped bubbles of gas
(mainly CO_2 from baking powder)

gas expands and batter rises

How could we explain this? Think microscopically! As the heat is applied to the gas, the gas molecules will gain energy and move much more rapidly. This rapid movement of molecules will result in more collisions of molecules with the walls of the balloon. These walls, being elastic, will expand resulting in greater volume.

If, however, the gas was trapped in a container made of glass or any other inflexible material, all volume will remain constant, but with increased temperature, the gas molecules will hit the walls more frequently resulting in higher pressure. The relationship between pressure and temperature if volume and number of moles are constant is expressed by Gay Lussac's Law. (P/T = constant)

Gay-Lussac's Law – Pressure/Temperature Relationship

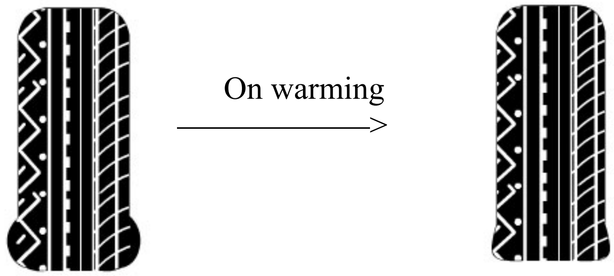
Pressure and temperature are directly proportional and this relationship is mathematically expressed in Gay-Lussac's Law as:

$$\frac{P}{T} = k$$

$$\frac{T_2}{T_1} = \frac{P_2}{P_1}$$ (V = constant, no. of gas molecules is constant)

This temperature/pressure relationship is observed in the automobile tires.

Why do automobile tires look flat on a cold winter morning?



As temperature rises during the day or while driving, the pressure goes up and the tires no longer look flat.

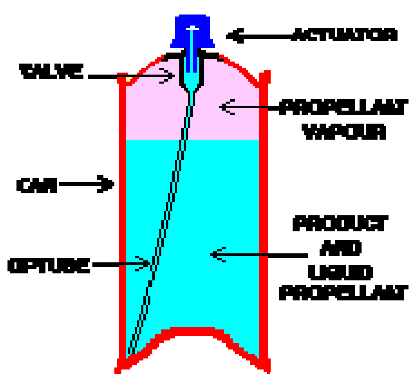
The same relationship explains the operation of aerosol cans.

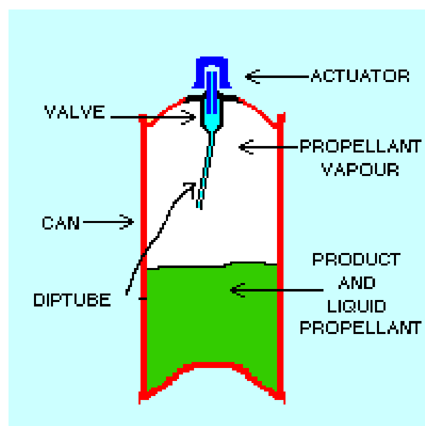
An "empty" aerosol can (deodorants, spray paint cans) is considered empty because it no longer delivers the contents. Actually the gas propellant pressure is equal to atmospheric pressure so no molecules flow from the can.

$$P_{propellant} = P_{atmosphere}$$

If the can is placed into a fire, the pressure in the can increases. The can eventually explodes because the pressure inside exceeds the strength of the soldered seam.

Aerosol Cans





Push top down
Propellant is propane
Can must be held upright to work

Spray paint or deodorant

$P_{propane} > P_{atm}$

Bend nozzle
Propellant is N_2O (laughing gas)
Can must be held upside down to work

Whipped cream

$P_{N_2O} > P_{atm}$

In both cans $P_{inside} > P_{outside}$, so gas will flow out of cans trying to equalize pressure; as the liquid is pushed from the can, the volume of gas increases and pressure drops ($V\uparrow$, $P\downarrow$).

“BALLOONS”

In 1980, in the early hours of a late October morning, in Denver, ***Julian Nott and his hot air balloon, “Innovation,” began the*** *record-breaking attempt at balloon elevation. The world’s top balloonists are now turning their attention to the circumnavigation of the earth.*

In June, two Soviet spacecraft, Vega 1 and Vega 2, bound for a 1986 rendezvous with Halley’s comet, released two helium balloons into Venus’ cloudy and mysterious atmosphere. The balloons, each ten feet in diameter, floated 33 miles above the planet’s dark side for nearly two days before they drifted around to its sunny side and expired. In their short life span they sent signals to earth about Venus’ weather. And while the balloons drifted above the planet, two Landers, also dropped by the Soviet spacecraft, probed its surface. Although the Landers transmitted for only about 20 minutes each, they were able to gather and analyze samples of the basaltic soil in that time.

Discover, August 1985

We all know that balloons filled with helium or hot air when released will rise in the air. We explained why balloons expand in a hot car and shrink in the winter outside. Why do hot balloons and helium balloons go up in the air? This is due to the density of gases and the effect of temperature on the density.

The density of a substance is an interesting physical property:

$$D = \frac{\text{mass}}{\text{volume}} = \frac{m}{V}$$

For a gaseous substance, the gas density is related to physical conditions such as pressure, temperature and molecular weight. It can be found by the following equation:

$$D = \frac{P\,(MM)}{R\,T}$$

Where P is pressure in atmospheres, T is Kelvin temperature, R is a constant (0.0821 L·atm/mol·K) and MM is the molar mass of the gas. If you study the gas density equation you can explain some interesting facts about ballooning (keeping in mind that less dense things float on top of more dense things).

Hot Air Balloons

*As T increases, D decreases

Gases are less dense at high temperatures:

*$D_{\text{hot air}} < D_{\text{cold air}}$

A hot air balloon will rise because of the above relationship. You can demonstrate this by holding a plastic bag over a hair dryer.

*As MM increases, D increases

Helium Balloons

Consider the following gases:

Helium	MM	= 4 g/mole
Air	MM	= 29 g/mole
Propane	MM	= 44 g/mole

Helium has a lower molar mass than air hence a helium balloon will rise if released into the atmosphere. That's why you need to hold on tight to those mylar helium party balloons! Weather balloons are helium balloons and are baggy at ground level. When released into the atmosphere, they expand ($P\downarrow$ so $V\uparrow$) due to the pressure decrease as altitude increases. Propane is denser than air, hence it sinks, and can be hazardous when used in households because it accumulates on the floor (can be ignited by refrigerator motors!).

Graham's Law of Effusion

If a skunk came into a large room and released its smell, we would have to evacuate the room. If a person drops ammonia, NH_3 bottle at the entrance to a large area even people on the other end will start crying. When a truck overturns and spills hydrochloric acid, HCl, the hydrogen chloride gas spreads and the neighborhood must be evacuated. In these examples, the gas molecules spread through the air moving through the space between the gaseous molecules, i.e., they DIFFUSED through the air. There is no barrier to their motion. The rate at which different gases diffuse depends on their molecular weight. The mathematical expression of this relationship is **GRAHAM'S LAW**. The lower the molecular weight of a gas, the faster it diffuses. (In actuality, the law describes gases that move through a small hole in a barrier. This process is called EFFUSION. However, the variables behave the same way).

Consider the person standing in the middle of a room. The following gases are simultaneously released at each end of the room:

Tear Gas
$C_6H_{11}OBr$
MM = 179 g/mole

Laughing gas
N_2O
MM = 44 g/mole

Is our person going to cry or laugh first? Well, which gas diffuses faster?

Shee will laugh first, because nitrous oxide, N_2O, has a lower molar mass and diffuses faster than tear gas. The relationship between molar mass and the velocity shows that the square root of the molar mass of a gas is indirectly proportional to the velocity.

Mathematically expressed this is:

$$\frac{V_A}{V_B} = \sqrt{\frac{MM_B}{MM_A}}$$

V_A, V_B = velocities or speed of molecules A and B respectively

or

$$\frac{d_A}{d_B} = \sqrt{\frac{MM_B}{MM_A}}$$

d_A, d_B = distance that the molecules A and B respectively covered

More Gas Laws…

The laws we have looked at so far generally have only two variables, but in real situations it is common for three variables or more to change. To accommodate that there are two more gas laws to discuss.

The <u>combined gas law</u> is used when pressure, volume, and temperature change. The form is:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ where 1 = one set of conditions and 2 = a second set}$$

The other law is the <u>Ideal Gas Law</u> used when there are several variables but only one set of each. The Ideal Gas Law is:

$$PV = nRT \text{ where R = gas constant } (0.0821 \ \frac{\text{L atm}}{\text{mol K}})$$

C. Misconceptions

1. Gases are not matter because they don't weigh anything. Changing something into a gas is a destruction of matter because you can't see it anymore. Gases, of course, do have mass. This can be demonstrated by weighing an evacuated container and then weighing it again when it is filled with air. Or it may be easier to weigh a bic lighter, release some of the butane gas; then, reweigh it, demonstrating the loss of mass.

2. When a gas is formed as a result of a chemical or physical change, matter has <u>not</u> been destroyed, only changed.

D. Questions before the Activity

Warm-Up Exercises

1. What happens to a basketball when you take it outside in the winter and bounce it? (Charles' Law)

2. Ask students how they could make measurements of volume and temperature. Use a balloon as an example.

3. Ask what they know about balloons in hot and cold weather and why

4. Ask what happens when they let go of a helium party balloon and why they think so

5. How would you prove that a gas is matter?

2. What happens to the bicycle tires in winter? Why? How does temperature effect volume?

3. Can you squeeze a balloon and make it smaller? Can you do that if it is filled with water?

4. Why do they keep gases like oxygen in hospitals in such heavy tanks?

Some Practice Questions About Gases

1. Predict the results of the following experiments. State exactly what you expect to see and explain the reasons behind your prediction.

a) A slightly inflated balloon is placed inside a bell jar that is resting on a vacuum table which is attached by way of a rubber tube to a vacuum pump. The pump is turned on and the air is removed from the region surrounding the balloon inside the jar.

b) The plunger of a hypodermic syringe is drawn upward and while in that position, the open end of the syringe is sealed off. Is the effort to push the plunger forward (downward) met by some resistance?

c) A "pulse glass" apparatus consists of a sealed tube with two bulbous ends. Inside the tube is a colored, low-boiling liquid under reduced pressure. One of the bulbous ends is now held in the hand for a couple minutes. What happens to the colored liquid?

d) A heavy-walled metal can is connected to a vacuum pump. The vacuum pump is turned on. What happens to the can? If a thin-walled container is used instead, what happens?

2. Consider the compressibility of gases. When you compress a gas, are you actually compressing the gas molecules? Explain your answer.

3. Why does a deep sea diver develop the "bends" (nitrogen gas bubbles accumulate in the blood) if he ascends too quickly? What disaster will occur if a submarine travels too deep into the ocean?

4. Why are we warned to keep aerosol cans away from high temperatures even if they are empty?

5. Identify the five gases below whose letters are scrambled.

MCABOXONODERIN ______________________________

NONE ______________________________

GNOYXE ______________________________

RGEHDNOY ______________________________

DESUTIRNOXIO ______________________________

For the gases you unscrambled above:

The lightest of these five gases is ______________________

The heaviest of these five gases is ______________________

One of these five gases is poisonous ______________________

One of these five gases was used by dentists ______________________

One of these gases is necessary for life ______________________

E. GLOSSARY

Atmosphere	A unit to express pressure; "standard pressure" is 1 atm = 760 mmHg = 760 torr = 76 cmHg
Boyle's Law	Pressure-volume relationship; temperature and the number of gas molecules are constant $P_1V_1 = P_2V_2 = P_3V_3 = \text{constant}$
Celsius (^oC)	Internationally used scale for measuring temperature, in which 100^oC is the boiling point of water at sea level and 1 atmosphere, and 0^oC is the is the freezing point. A temperature given in Celsius degrees may be converted to the corresponding Fahrenheit temperature by multiplying it by 9/4 (or 1.8), and adding 32
Changes of state	a change from one physical state to another, for example, from the solid to the liquid state, or from the liquid to the gaseous state, or the reverse; when a change of state occurs, the chemical substance remains the same; only its physical state changes
Charles' Law	Temperature-pressure relationship; volume and the number of molecules are constant $\frac{V_1}{T_1} = \frac{V_2}{T_2} = \frac{V_3}{T_3} = \text{constant}$
Condensation	a change of state from the gaseous state to the liquid state
Diffusion	The spontaneous mixing of one substance with another when in contact; diffusion occurs most readily in gases, less so in liquids, and least in solids
Dipole-dipole interactions	Type of intermolecular force in which opposite ends of polar molecules are attracted to one another
Distillate	The product of distillation
Distillation	A separation process in which a liquid is converted to vapor and the vapor is then condensed to a liquid in a different container
Electrostatic force	Attraction between positive ions (cations) and negative ions (anions); this strong force results in the solid state of ionic compounds
Evaporation	The change of a substance from the liquid to the gaseous or vapor phase

Fahrenheit	A temperature scale in USA in which melting point of water is 32^oF and boiling point of water is 212^oF
Freezing	the physical change that occurs when a substance changes from the liquid to the solid state
Graham's Law of Effusion	Smaller molecules diffuse faster than large molecules $$\frac{\text{Rate of diffusion } (v_1)}{\text{Rate of diffusion } (v_2)} = \sqrt{\frac{MM_2}{MM_1}}$$
Heat of fusion	the energy required to change a substance from the solid state to the liquid state
Heat of vaporization	the energy required to change a substance from the liquid state to the gaseous state
Hydrogen bonding	Special IMF that occurs when a H that is bonded directly to an O, N, or F, is attached to an O, N or F on a nearby molecule
Ice	water under the solid state
Intermolecular forces (IMF's)	Attractive forces between molecules which are responsible for keeping matter in the solid or liquid state
Kelvin	K; the temperature scale used in gas laws; the absolute is 0 K; melting point of water is 273K; boiling point of water is 373K
London forces	Weak type of intermolecular force caused by temporary distortion of electron clouds; only IMF between non-polar covalent molecules; also called dispersion forces
Melting	the physical change associated with a transition from the solid state to the liquid state
Milli	Prefix meaning 10^{-3} unit or $1/1000^{th}$ part
Milligram	(mg) is one-thousandth gram or 10^{-3} g
Milliliter	mL is one-thousandth liter or 10^{-3} L; it is also the volume occupied by one gram of pure water at 4^oC and 760 mm Hg; 1 mL = 1 cubic centimeter

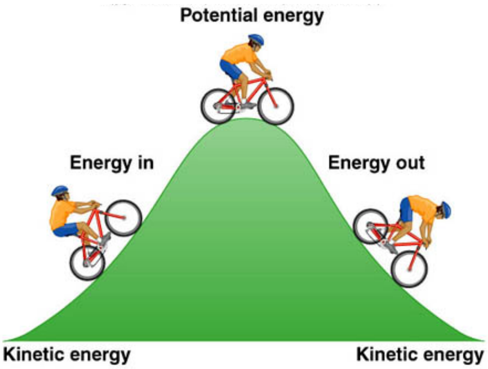
Mole	Amount of any substance which contains 6.023×10^{23} units of things; mol is the abbreviation
Physical state	the solid state, the liquid state, or the gaseous state
Pressure	Force per unit area
Steam	water under the gaseous state
STP	Standard temperature and pressure; 273K, 760 mm Hg
Sublimation	A direct conversion of a substance from solid to vapor without appearing in the intermediate liquid state; examples are solid carbon dioxide, naphthalene (moth balls), iodine
Temperature	Degree of "hotness" – measured in degrees, ^{o}C, ^{o}F, K; it can be described as a measure of the direction of the flow of heat as heat travels from areas of higher temperature to areas of lower temperature
Vaporization	the physical change from the liquid to the gaseous state
Volatile liquid	A liquid, usually an organic solvent, that has high vapor pressure at room temperature and evaporates readily; it has weak IMF's
Volume	Space occupied by matter; is measured in liters, L or milliliters, mL, cubic centimeters, cm^3

CHAPTER 6

Thermal Energy

B. Background

Matter has been defined as anything that has mass and takes up space. Energy is not matter. But matter and energy routinely interact with one another. Energy is the capacity to do work. There are two basic types of energy – kinetic energy and potential energy. Kinetic energy is the energy something possesses because of its motion; potential energy is the energy something possesses because of its position. Energy can be converted from potential to kinetic or vice versa. In a hydroelectric power plant, the potential energy of the water at the top is converted to kinetic energy as the water falls down the incline. The kinetic energy of the water is then converted to mechanical energy which turns the turbines and eventually converts mechanical energy to electricity. In all this, energy is neither created nor destroyed. Instead it is converted from one form to another.



Thermal Energy is the internal energy of an object. Every object has tiny particles in it which are in motion. The hotter the object the more thermal energy it has. When the temperature is high, the small particles will be moving faster due to the increased heat.

Temperature and Thermal Motion

All matter is composed of particles in constant motion. They therefore have kinetic energy. Temperature is directly proportional to the average kinetic energy of particles. Two objects are at the same temperature when the average kinetic energy of the particles in each is the same. Thermal energy is the total of the energy of all particles in the object. The larger the energy mass, the more particles, and the greater the thermal energy.

Heat is the energy that flows from an object of high temperature to one of low temperature. If ice cubes are left in a glass at room temperature, the heat will flow from the air at room temperature into the ice causing the ice to melt and the resulting water to rise in temperature. Heat is similar to work in that both are energy being transferred. Whereas, heat is the energy transferred between objects at different temperatures, work is the energy transferred when one object exerts an unbalanced force on another object. Heat and work are measured in SI units called joules (J). The calorie is another unit of heat energy (not an SI unit). One calorie is the amount of heat necessary to raise 1 gram of water by 1^oC (1 call = 4.184 J). When we speak of the "calories" in food, we are actually referring to kilocalories or 1,000 calories. Therefore, a slice of white bread that has 55 Calories (note the large C) contains enough heat energy from that bread (5500 calories, small c) to raise 5,500 g of water 1^oC or 100 g of water 55^oC.

Thermal Energy Transfer

Thermal energy is transferred in three ways: conduction, convection, and radiation.

Conduction involves the transfer of energy through matter from particle to particle. Usually conduction is more effective in solids than in liquids or gases. Can you explain why? Metals are good conductors of heat. Diamond, which is a form of carbon and a non-metal, happens to be the best known thermal conductor.

Convection is the transfer of heat through fluids by the movement of matter. Liquids and gases are called fluids because they flow. Wind and ocean currents are fluids that transfer heat by convection.

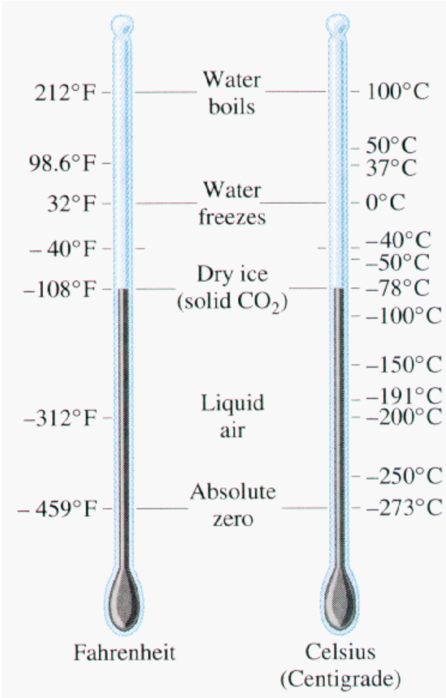
Radiation is the transfer of energy that does not require the presence of matter. French fries are frequently kept hot in restaurants using heat lamps, yet the heat lamp never touches the food. The heat is transferred by radiation.

Heat flows from areas of high temperature to areas of low temperature. Insulation is used to reduce the flow of heat. Plastic foam, fiber glass, and cork are all examples of insulators. Air is a poor conductor of heat, so it acts to prevent the flow of heat in double pane windows.

Temperature Measurement

The common device for measuring temperature is a thermometer. Most laboratory thermometers are composed of a heat-expanding liquid such as mercury or ethanol trapped in a sealed glass tube. The tube is inscribed with values based on standards and the expansion characteristic of the liquid.

There are two common temperature scales in use in our everyday lives. The Fahrenheit scale is a non-metric scale that sets the freezing point of water as 32^oF and the boiling point as 212^oF. This scale is commonly used in the United States for cooking and weather information. The scale used in lab and most other countries in the world is the metric system Celsius scale (formerly called the centrigrade scale). On the Celsius scale water freezes at 0^oC and boils at 100^oC.



Here are some conversions between these two scales.

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

Specific Heat

Specific heat (C_p) is the thermal energy required to cause a change in temperature of a substance by one Celsius degree. The units for specific heat are joules per gram Celsius degree (J/g· °C). We know that different substances have different specific heats. If a silver spoon is placed in a dish of hot food, the spoon will get hot very quickly. If a stainless steel or plastic spoon is placed in the same dish of hot food, it will have a much smaller change in temperature.

Another example which illustrates differences of specific heat is the temperature change which occurs in the ocean water compared to the temperature change which occurs in sand, both under the same ninety degree temperature. The sand will get hot; the water only warms slightly. This is in accord with the difference in specific heats. The specific heat of seawater is 3.90 J/g· °C whereas the specific heat of sand, $S1O_2$, is 0.80 J/g· °C.

Examples of the Specific Heat of Selected Substances

Substance	**Specific heat J/g °C**	**Substance**	**Specific heat J/g °C**
Aluminum	0.90	Ice	2.10
Iron/steel	0.45	Wood	1.70
Copper	0.39	Nylon	1.70
Brass	0.38	Rubber	1.70
Zinc	0.38	Marble	0.80
Silver	0.23	Concrete	0.85
Mercury	0.14	Granite	0.84
Tungsten	0.135	Sand	0.80
Platinum	0.13	Glass	0.67
Lead	0.13	Carbon	0.50
Hydrogen	14.0	Ethanol	2.40
Air	0.72	Paraffin	2.10
Nitrogen	1.04	Water	4.186
Steam	2.00	Seawater	3.90

How to Measure Heat Flow

It is possible to measure the change in thermal energy of a given material by measuring the mass of the material and its initial temperature. The final temperature is recorded after the material is heated or cooled. In order to calculate the change in thermal energy, it is also necessary to know the specific heat of the material.

The change in thermal energy (Q) equals the temperature change multiplied by the mass and the specific heat.

$$Q = c \times \Delta T \times m$$

where c = specific heat, ΔT = change in temperature, m = mass

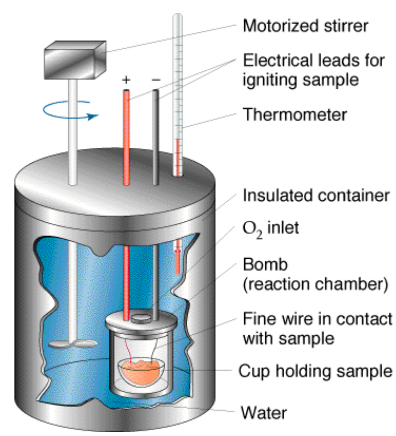
The value of ΔT is always positive regardless of whether heat is lost or gained.

When heat is gained,
$\Delta T = T_{final} - T_{initial}$

When heat is lost,
$\Delta T = T_{initial} - T_{final}$

Calorimeters

Calorimeters are instruments used to measure changes in thermal energy. Frequently calorimeters have two chambers. In the inside chamber a reaction occurs. The outer chamber contains water. When the reaction occurs, heat is given off. It heats the water in the outer chamber. In an ideal calorimeter, the heat given off in the inner chamber will be equal to the heat absorbed by the water in the outer chamber. Usually the outer walls of a calorimeter are insulated like a thermos bottle. The heat released by the reaction is equal to the heat gained by the water. Again, we use the equation:



$$Q = c \times \Delta T \times m$$

but in a slightly different way. The following calculations illustrate this idea.

Problem 1. A sample of food is burned in a calorimeter. The original temperature of the food and the water in the outer chamber of the calorimeter is 20 °C. After the food is burned, the temperature of 100 g of water has risen to 35 °C. How much heat is transferred by burning the food?

Basic Equation: $Q = c \times \Delta T \times m$

Given:

Initial Temperature	=	20^{o} C	Unknown: Q
Final Temperature	=	35 ^{o}C	
Mass	=	100 g	
c	=	4.180 J/g$\cdot$ ^{o}C	

Solution

$$Q = c \times \Delta T \times m$$

$$Q = 4.18\ \frac{J}{g\ ^{\circ}C} \times (35\ ^{\circ}C - 20\ ^{\circ}C) \times 100\ g$$

$$Q = 6270\ J = 1499\ cal\ \text{ or } 1.5\ kcal\ (2\ \text{sig figs})$$

Problem 2. A 40 g piece of carbon is cooled from 80 ^{o}C to 25 ^{o}C. How much heat is lost by the carbon?

Basic Equation: $Q = c \times \Delta T \times m$

Given:

Initial Temperature	=	80 C^{o}	Unknown	=	Q
Final Temperature	=	25 C^{o}			
Mass	=	40 g			
c	=	0.710 J/g$\cdot$ ^{o}C			

Solution:

$$Q = c \times \Delta T \times m$$

$$Q = 0.710\ \frac{J}{g\ ^{\circ}C} \times (80\ ^{\circ}C - 25\ ^{\circ}C) \times 40\ g$$

$$Q = 1562\ J = 373\ cal$$

C. Misconceptions

1. Many students think that when something is heated, it weighs more because they view heat as having physical substance. You cannot see the heat that would raise the temperature of a liquid by 10 degrees – can only detect it by using a thermometer or noting that it feels hotter by touching it. You cannot see air either. Of course the difference is that air is matter and heat is a form of energy. This difference can be vividly demonstrated to students by weighing both cold and warm water on a platform balance.

2. The terms temperature and heat are often confused. Temperature is a measure of hotness or coldness. If the temperature of an object is higher than its surroundings, the heat will flow from the object to the surroundings – or vice versa. A large vat of molten iron and a small crucible of molten iron may both be at the same temperature – but the total amount

of heat in the large vat is greater. Quantities of heat are measured in joules or calories. Temperature is measured by degrees Fahrenheit, Celsius, or Kelvin.

D. Warm-up Exercises

1. What kinds of energy are found in your classroom?

   Answer: Electricity, heat (what source), chemical in food, batteries, are all possibilities. There are probably more.

2. What is the fundamental source of all our energy?

3. What is the name given for the conversion of solar energy into plant material? Can people use the sun in the same way as plants?

4. What do you think has more mass, a hot object or a cold object? Why?

Additional Questions and Activities after the Lab

1. How does a change in thermal energy affect the motion of particles in an object?

2. Place a cup of sand, a cup of water, a cup of Styrofoam peanuts in a sunny window. Compare the temperature of each after 30 minutes. Are they the same? Why or why not

3. When you hammer a nail, why does the nail head get hot?

4. Which has more potential energy, a 150 pound diver on the high dive or a 150 pound diver on the spring board? Why? Compare their potential energies when they both are in the pool!

5. A carton of milk is placed next to a carton of orange juice in the refrigerator. Both are left there overnight. Does the milk have the same average kinetic energy as the orange juice? Explain your answer.

6. A 0.04 kg sample of aluminum is cooled from $80^{o}C$. How much heat is lost by the aluminum? The specific heat of aluminum is $0.920\ J/g\cdot{}^{o}C$.

E. Glossary

Calorie	the amount of heat energy necessary to raise 1 g of water by $1^{o}C$
Calorimeter	instrument used to measure changes in thermal energy
Conduction	a method of thermal energy transfer whereby energy is transferred through matter from particle to particle; conduction is usually more effective in solids than in liquids or gases
Convection	a method of thermal energy transfer whereby energy is transferred through fluids by the movement of matter
Endothermic process	a process in which energy is absorbed from the environment
Exothermic process	a process in which energy is evolved to the environment
Expansion	an increase in volume; as gas expands when its temperature is increased; most solids and liquids also expand on increase in temperature
Hydroelectric power	electricity generation from the energy of falling water
Insulation	materials added to impede the flow of heat
Joule	a unit to measure energy; 4.184 J = 1 cal
Kinetic energy	the energy associated with motion; kinetic energy depends upon the mass and velocity of a substance
Mechanical energy	energy associated with the movement of parts of machines
Potential energy	energy associated with the position of a substance; potential energy depends upon the mass, the force of gravity and the height of a substance
Radiation	the transfer of energy that does not require the presence of matter
Sand	the chemical, silicon dioxide
Specific heat	the energy required to change the temperature of a kilogram of a substance 1 degree Celsius ($1^{o}C$)
Temperature	a measure of the average kinetic energy of particles
Thermal energy	the internal energy of an object; it is equal to the total energy of all particles in an object

Work	the energy transferred when one object exerts an unbalanced force on another object

CHAPTER 7

Solutions

B. Background

What do fruit drinks, coffee, tea, vinegar, and sea water have in common? They are solutions. Solutions are homogeneous mixtures with two or more components. Mixtures are far more common than pure substances. True solutions are homogeneous mixtures containing

particles the size of molecules. The majority of chemical reactions occur in solution because the particles are moving from place to place and reactants can come into contact with one another more easily. Solutions and other liquid mixtures, colloids and suspensions, should not be mistaken as identical. Filtered tea or kool-aid drink do not settle even on standing – they are true solutions. Parts of chocolate drink and orange juice will settle after standing for a short time – they are suspensions. Light will pass through tea or any other true solution without being scattered, whereas light passed through a colloid (such as a dilute starch and water mixture or skim milk) or a suspension is scattered by the suspended particles. The major, although not obvious, difference among solutions, colloids, and suspensions is the size of the mixed particles.

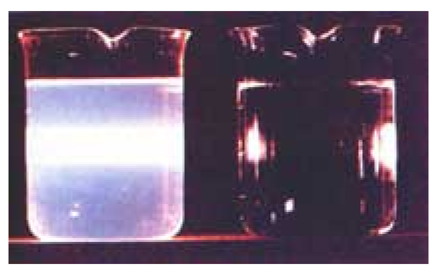
The characteristic properties of solutions are:

1. Solutions are homogeneous mixtures; they appear to be uniform throughout and have the same composition throughout;

2. Solute particles are small; they are usually individual molecules or ions moving about at random;

3. A solution is a mixture rather than a compound; the proportions of solute and solvent can vary over a wide range of values;

4. The particles in true solutions do not settle out or deflect a beam of light.



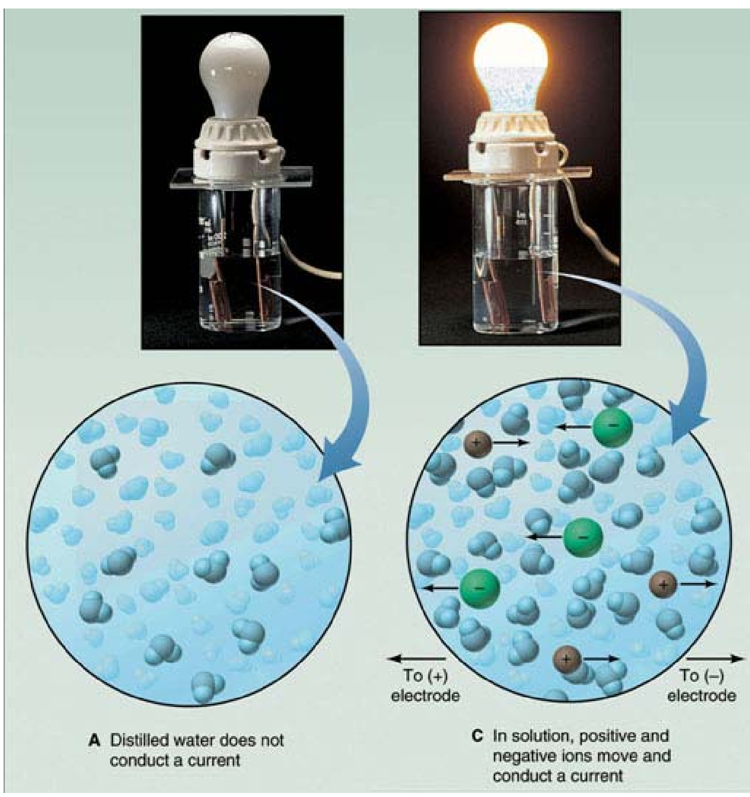
Solutions are mixtures which consist of a SOLVENT and at least one SOLUTE. The solute is usually the component present in lesser amount. The solute dissolves in the solvent to produce a solution. The solvent is a part of the mixture that is typically present in the larger amount and is responsible for dispersing the solute particles. If water is one of the components of the mixture, it is usually considered the solvent.

Although we usually think of solutions as being in the liquid state, the physical states of solute and solvent can be any combination of states and are solutions as long as they meet the particle size requirement. Some examples are given below.

SOLUTE	Gas	Gas	Liquid	Solid	Solid
SOLVENT	Gas	Liquid	Liquid	Liquid	Solid
EXAMPLE	Air	Carbonated Beverages	Alcoholic Beverages Gasoline	Sea Water	Alloys like Sterling Silver or Steel

The most common solutions are liquid-liquid and solid-liquid mixtures. The most common solvent is water. We refer to water solutions as <u>aqueous</u> solutions.

The particles of true solutions can be molecules or ions. Solutions of sugar in water, carbon dioxide in water, alcohol in water, coffee, or tea in water all contain undissociated solute <u>molecules</u> held together by covalent bonding. Due to the absence of charged particles (molecules are neutral) these solutions do not conduct electricity. They are <u>non-electrolytes</u>. Solutions of table salt, NaCl in water, seawater (that contains a number of other salts in addition to NaCl), hard water (which contains ionic compounds of iron, calcium and/or magnesium), and NaOH (lye) in water which are ionic compounds, contain <u>ions</u> in water. Due to the presence of ions or charged particles these solutions are <u>electrolytes</u>, since they conduct electricity.



NON-ELECTROLYTES
Do not conduct electricity
Examples: CO_2, sugar

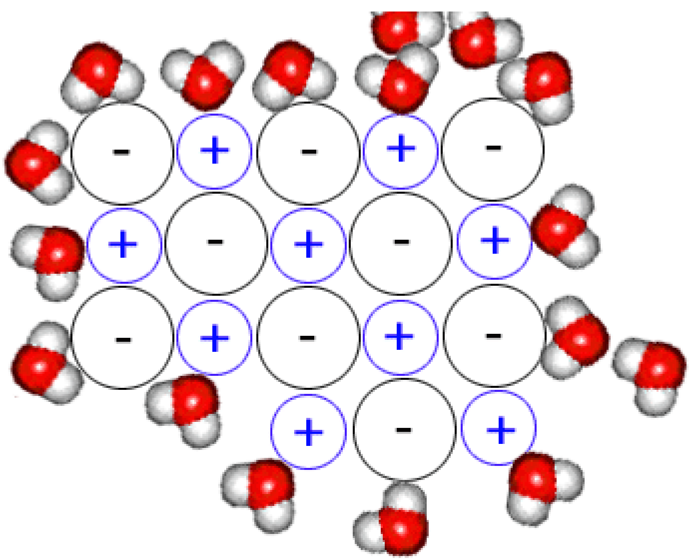
ELECTROLYTES
Conduct electricity
Examples: NaCl, $NaHCO_3$

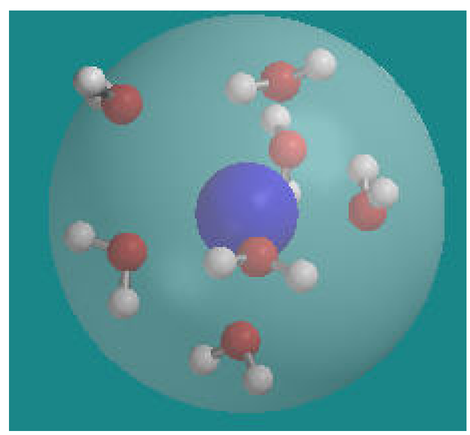
Strong acids interact with water to produce ions in solution (through ionization). Strong bases and salts are completely dissociated in aqueous (water) solutions (the ions that are already present in the compound are separated) thus producing solutions that are strong electrolytes or good conductors of electricity. Weak acids that are only partially ionized in aqueous solutions are not as good conductors of electricity and produce solutions that are weak electrolytes. Solutes that do not dissociate or ionize when dissolved in water, like carbon dioxide, sugars, alcohol, produce solutions that do not conduct electricity and are nonelectrolytes.

The Solution Process

1. IONIC COMPOUNDS. We already know that ions are present in a solution of salt water. The dry salt crystal consists of a regular arrangement of interlocking positive sodium and negative chloride ions. During dissolving, these ions are separated from one another and are released into the solvent. Since oppositely charged ions are being separated, energy is required. Where does this energy come from? Remember that bond making is energy releasing! When a salt crystal is dropped into the polar covalent solvent water, the positive ends (poles) of the water molecules are attracted to the negative chloride ions, and the negative ends of the water molecules are attracted to the positive sodium ions at the surface of the crystal. This attraction releases sufficient energy to overcome the electrostatic force holding the crystal together.

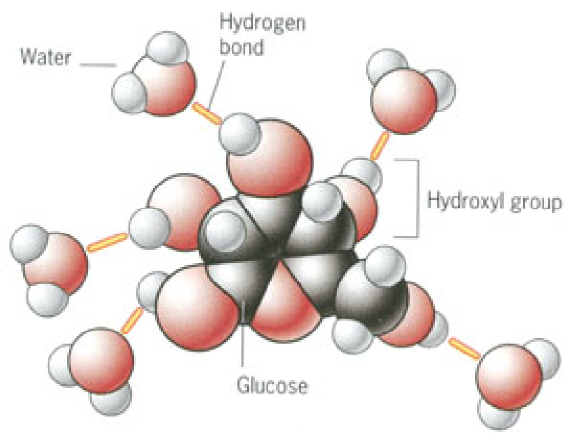
   Furthermore, once the ions are separated, or dissociate from the crystal, they are prevented from re-joining each other because many water molecules cluster around the ion, with their oppositely charged poles all pointing toward the ions.

   If the energy released by the attraction between the solvent and the solute is greater than the energy needed to dissociate the ions, the solution process is exothermic – and the beaker will feel hot as the solute dissolves. If the beaker feels cold then the energy needed to dissociate was greater than that provided by the action of the solvent, and heat was taken from the surroundings (endothermic). In cases where the energy holding a crystal together (referred to as LATTICE energy) is very high, the salt will be insoluble in water. If a salt dissolves in water, and the beaker feels neither hot nor cold, the energy provided by the action of water (HYDRATION ENERGY) is equal to lattice energy and the solution is referred to as an IDEAL SOLUTION.

2. COVALENT COMPOUNDS. Sugar (sucrose or table sugar) is a polar covalent compound. In solid sugar, the "crystal" consists of a regular arrangement of molecules held together by the attraction of the oppositely charged poles (DIPOLE-DIPOLE INTERATIONS). When placed in water, hydration energy is supplied by the attraction of the polar water molecules to the sugar molecules. This energy is sufficient to overcome the lattice energy of the sugar crystal; the molecules are surrounded by water molecules and are "in solution." The picture here shows a molecule of glucose which is a component of table sugar.



If a non-polar covalent solute, such as octane, C_8H_{18},, is placed in water, there is nothing the water molecules can be attracted to because there are no positive or negative poles. The liquid octane is held in the liquid state only by the weak LONDON forces. Consequently, octane, and other non-polar compounds such as vegetable oil do not dissolve in water.

Concentration of Solutions

We deal with solution concentrations in our daily life when we talk about strong and weak coffee. Strong coffee contains more coffee, more caffeine, more of all solutes; it is more concentrated. Concentration describes how much solute is in a given amount of solution. Concentrated solution is a semi-quantitative term implying presence of a high quantity of solute. Dilute solution implies low amount of solute. There are other precise ways of expressing concentrations numerically. Vinegar is "5 percent acetic acid" (5 grams of acetic acid in 100 grams of vinegar), rubbing alcohol labels indicate "70 percent isopropyl alcohol by volume" (70 mL of isopropyl alcohol in 100 mL of solutions). Percentage by weight or volume is only one way of expressing solution concentration.

$$\% \text{ Solution} = \frac{\text{amount of solute}}{\text{amount of solution}} \times 100$$

amount can be expressed in grams (g) or milliliters (mL)

There are other ways, such as molality (M) – moles of solute in 1 L of solution, molality (m) – moles of solute in 1000 g of solvent.

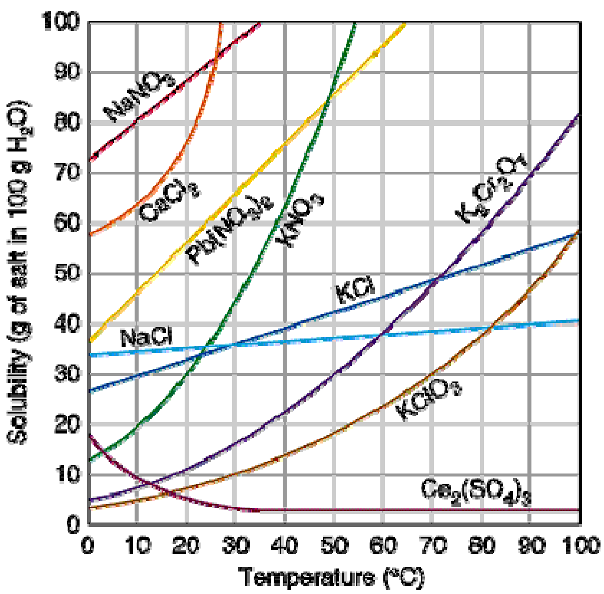
$$\text{Molarity} = \frac{\text{moles of solute}}{\text{1 liter of solution}} \qquad \text{Molality} = \frac{\text{moles of solute}}{\text{1 kilogram of solvent}}$$

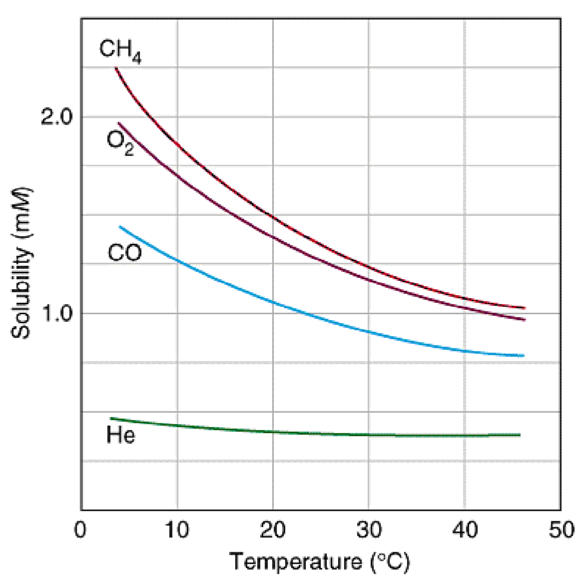
The maximum amount of solute that can dissolve in a given amount of solvent is referred to as the solubility and is different for each compound. The solubility of table salt (NaC1) is not the same as that of sugar ($C_{12}H_{22}O_{11}$), sodium hydroxide (NaOH), or calcium acetate $Ca(C_2H_3O_2)_2$. In addition, the solubilities for each compound are different at different temperatures.

Solubility of Some Solids in Water

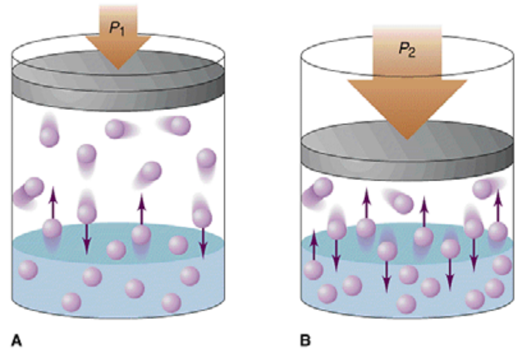
NAME	FORMULA	SOLUBILITY grams/100 mL H_2O	
		0°C	**100°C**
sodium chloride (table salt)	NaCl	35.7	39.12
sodium hydroxide (lye)	NaOH	42	347
magnesium hydroxide (Milk of Magnesia)	$Mg(OH)_2$	0.0008	0.004
cerium (III) sulfate	$Ce_2(SO_4)_3$	18	2.5
table sugar (sucrose)	$C_{12}H_{22}O_{11}$	100	500

The solubility of $Mg(OH)_2$ is very low at 0°C and increases five times at 100°C (but is still very low!); solubility of NaOH is much higher at 0°C and increases more than 3 times at 100°C. Most solubilities increase with increase in temperature. However, there are compounds such as cerium (III) sulfate ($Ce_2(SO_4)_3$) that have a lower solubility at higher temperatures. In general, the solubility of gases decreases with an increase in temperature. Below are graphs of solubility versus temperature for solids (left) and gases (right).





Pressure does not affect the solubility of liquids or solids but has a big effect on the solubility of gases. When pressure increases, gases will dissolve more. This relationship is referred to as Henry's Law. Carbonated beverages retain their bubbles in the bottle or can because of the pressure under the top. When a diver goes to greater water depths, more nitrogen dissolves in his bloodstream. If he comes to the surface too fast the gas undissolves and forms bubbles that results in a serious condition known as the bends.



A solution that has all the dissolved solute that it can possibly dissolve is called a <u>saturated</u> solution. A solution containing less than the maximum possible amount of solute is termed <u>unsaturated</u>. Sometimes if a solution saturated at high temperature is cooled very slowly and there is no agitation along the way, a <u>supersaturated</u> solution can be produced which contains higher amounts of solute than its actual limit. A small physical agitation of such solution will result in immediate crystallization of solute. The supersaturated solution becomes saturated with excess solute precipitating out.

C. Misconceptions

1. Students may believe that a solute, such as salt "disappears" (in the sense that it is no longer there) when it dissolves in water. This can be dispelled by asking students if they can distinguish salt form pure water by taste or by boiling the water off and retrieving the salt. This is a problem of 'out of sight – out of mind' and also occurs when dealing with unseen gases.

2. Students may think that all liquids conduct electricity because of warnings about handling electrical appliances with wet hands. Having them test various liquids and liquid mixtures can dispel this notion. Be sure to include distilled water in the testing. Since it contains only minuscule amounts of H_3O^+ and OH^- it will not be a conductor. It may be wise to point out to students that even if they wet their hands with distilled water, materials from the skin would dissolve in the water and make it a conductor. So the wet-hands warning is always applicable.

D. Warm-Up Exercises

1. How can you get sugar to dissolve in your ice tea faster? (How can we increase the rate of dissolution of solute into a solvent?)

   This can then be tried out with students and they can time how long it takes to dissolve a given amount of sugar in water using different techniques.

Answer:

Three ways to increase the rate:

1) Stir or mix – get the spoon going
2) Grind the solute – powered sugar > granular sugar > sugar cubes
   Note: In order for dissolving to occur the solvent must be in contact with the solute
3) Increase the temperature – put the sugar in the hot tea before adding the ice

2. Give examples of solutions from daily life. How do you know they are solutions?

3. Is seawater a solution? How would you prove with a simple experiment whether it is pure water or a solution?

4. "I want my Kool-Aid with just the right amount of sweetness as my mom makes. How can I be sure to make it the same?"

Follow-Up Exercises

1. Using data from a solubility table, select various amounts of a given solid. Draw labeled pictures of what beaker, all containing the same amount of water, but with the different amounts of solutes added would look like.

2. Perform the light test for solutions. (This can also be done as an introduction – to indicate that two clear liquid mixtures may not be true solutions. Clear colloids will scatter light (with their larger particles) whereas true solutions will not.

   (A clear colloid can be made by adding a few drops of dilute acid to an aqueous solution of $Na_2S_2O_3$ (sodium thiosulfate). Sulfur particles aggregate to form particles large enough to scatter the light.) The scattering of light is referred to as the Tyndall Effect. Place the liquid mixtures in two clean beakers and aim a flashlight or projection lamp at each. Observe the beakers from the side. The one that appears brighter from the side is the colloid.

Additional Questions

1. Suppose you were handed a solution and asked if it is a true solution, colloid, or suspension. How can you tell which it is?

2. Classify the following as a solution, colloid or suspension and explain why: milk, Phillips Milk of Magnesia, cherry Kool-aid, orange juice with pulp, French salad dressing, Listerine mouthwash, bottled water, hot chocolate

3. What is the difference between a concentrated and a dilute solution? Give an example of each. Why might the designations concentrated or dilute be inappropriate to use to some situations?

4. A bottle in a drug store contains a label "3 percent hydrogen peroxide." What does it mean?

5. Can a solution be both saturated and dilute?

6. How could you determine the concentration of sugar in a can of soda?

Demonstrations

1. Dissolve some sugar (or salt) in a glass of water and allow the glass to sit and the water to evaporate. What do you find on the bottom of the glass? (Sugar (or salt) crystals).

2. a) In a zip lock bag, place a little solid calcium chloride and some water and close the bag. What do you observe? (What do you see and feel?)

   b) Repeat the same procedure using ammonium chloride.

   c) Repeat the same procedure using table salt.

   Compare the observations in a), b), and c).

   Is the solution process of these three solids exothermic or endothermic? or neither?

E. Glossary

Colloid	a non-true solution containing particles that are larger than molecules; Colloids can be clear or cloudy, cannot be separated by filtration, but particles may pass through cell membranes
Crystal	regular arrangement of ions or molecules in a solid

Endothermic	term describing the process or change that takes place with absorption of heat or energy; endothermic reactions need heat; an example is dissolving of ammonium chloride in water $NH_4Cl \xrightarrow{H_2O} NH_4Cl\ (aq)$
Exothermic	term describing process or chemical reaction which is accompanied by evolution of heat; heat is given off in exothermic reactions as one of the products; an example is the combustion of methane $CH_4 + O_2 \dashrightarrow CO_2 + 2H_2O + \text{heat}$
Hydration energy	energy provided by attraction of water to particles in a solute
Ideal solution	a solution in which hydration energy equals lattice energy
Immiscible	term used to describe a liquid that will not mix with another liquid
Insoluble	unable to dissolve in a certain solvent
Lattice energy	energy required to break-up a crystal
Miscible	term used to describe a liquid that will mix with another liquid
Precipitate	solid particles that settle out of a liquid mixture by gravity, usually as a result of a chemical reaction
Precipitation	a process where a substance precipitates (settles out) out from a liquid mixture by the action of a chemical reagent, electricity, or heat
Saturated solution	solution which contains as much solute as possible; no more solute can be dissolved at that temperature
Soluble	able to dissolve in a given solvent
Solute	substance that dissolves to make a solution
Solution	homogeneous mixture of solute in solvent
Solubility	maximum amount of solute that dissolves to make a saturated solution at a given temperature
Solvent	substance capable of dissolving another substance (solute)
Suspension	a mixture in which the particle size is so large that the solid settles on standing; the component parts of a suspension can be separated by filtration

CHAPTER 8

Acids and Bases

B. Background

Acids and bases are compounds that we encounter frequently in our daily life. We clean with ammonia and lye, two familiar bases. We use vinegar (acetic acid) in salads and cooking. We drink beverages made tart by citric acid and phosphoric acid. We take vitamin C (ascorbic acid). Lactic acid is formed in sour milk, overworked muscles, and is responsible for the sour taste of yogurt. We produce hydrochloric acid in our stomachs and treat the excess acid with antacid tablets that contain bases. Our bodies produce and consume acids and bases, maintaining a delicate balance necessary to our good health and well-being.

Acids

The English word "acid" comes from acidus, Latin for "sour." Acids are compounds that:

- Have a sour taste
- Are neutralized by bases in a neutralization reaction in which salt and water are produced
- Turn the indicator dye litmus from blue to red
- Produce hydronium ions (H_3O^+) in aqueous solutions
- Have a pH below 7.0
- React with active metals (such as zinc, iron, tin, magnesium) to dissolve the metal and produce hydrogen gas

Compounds that contain hydrogen and a non-metal (HCl), and H and negative polyatomic ions (HNO_3) are acids. There are other compounds that do not fit this picture that can also be considered acidic!

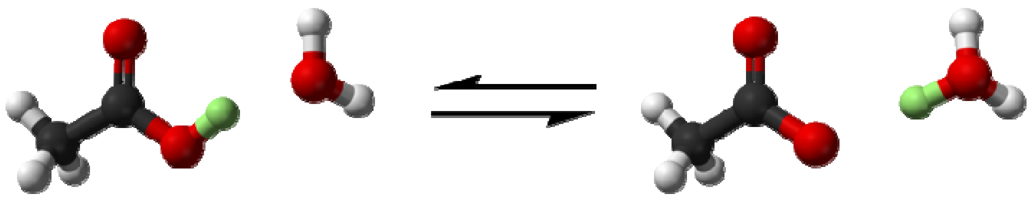
While we use and consume many acids in our daily life, there are those that are toxic or destructive: the leaves of rhubarb contain high concentrations of poisonous oxalic acid, $H_2C_2O_4$; concentrated hydrochloric acid, HCl, nitric acid, HNO_3 and sulfuric acid H_2SO_4 can cause severe burns and death.

Acids produce hydronium ions when dissolved in water. This process is known as <u>ionization</u>. A hydrogen ion, H^+ (a hydrogen atom minus its electron) from the acid attaches itself to a water molecule thus producing a <u>hydronium</u> <u>ion</u> (H_3O^+)

Acid + water ---> H_3O^+ + anion

The picture that follows shows the donation of a hydrogen ion from acetic acid to water.

CH_3COOH + H_2O -------> CH_3COO^- + H_3O^+



acetic acid (proton donor) | water | acetate ion | hydronium ion

$HCl + H_2O \text{ ---> } H_3O^+ + Cl^-$ (chloride anion)

$HNO_3 + H_2O \text{ ---> } H_3O^+ + NO_3^-$ (nitrate anion)

This characteristic of acids to ionize and produce or donate hydrogen ions causes acids to be called “PROTON DONORS”. A hydrogen ion is the proton since the ordinary hydrogen nucleus consists of only a single proton and no neutrons).

Different acids produce different concentrations of hydronium ions. Acids that produce high concentrations of hydronium ions and are completely ionized are strong acids. Those that produce low concentrations of hydronium ions are weak acids. In weak acids, only a small percentage of the molecules react with water to produce hydronium ions. The classification of acids as strong or weak is a measure of the degree of ionization.

Some Familiar Acids

Name	Formula	Classification
Sulfuric acid	H_2SO_4	Strong
Nitric acid	HNO_3	Strong
Hydrochloric acid	HC1	Strong
Phosphoric acid	H_3PO_4	Moderate
Hydrogen sulfate ion	$(HSO_4)^-$	Moderate
Lactic acid (from milk)	$CH_3CHOHCOOH$	Weak
Acetic acid (vinegar)	CH_3COOH, $H(C_2H_3O_2)$	Weak
Boric acid	H_3BO_3	Weak
Hydrocyanic acid	HCN	Weak
Citric acid (fruit)	$C_6H_8O_7$	Weak

Strong acids can cause serious damage to skin and flesh when they are concentrated. They cause holes in natural fibers such as cotton, silk, and wool. They destroy most synthetic fibers such as nylon, polyesters, and acrylics.

Concentrated and dilute solutions of acids should not be confused with strong and weak acid. Sulfuric acid is a strong acid. It can be concentrated (i.e., contain high percentage of H_2SO_4 in solution) or dilute (i.e., contain low percentage of H_2SO_4 in solution), but is strong in both cases. "Strong" and "weak" are terms that refer to the degree of ionization of the acid, or production of hydrogen ions and/or hydronium ions in aqueous solutions. Consequently, strong acids are good conductors of electricity (many ions are produced), while weak acids are poor conductors of electricity (there are few ions present in the solution).

Generally acids react with metals to dissolve the metal and produce hydrogen gas. This is not true with all metals, or all acids. Gold and platinum, for example, do not dissolve even in strong acids. Dilute solutions of weak acids react very slowly with most metals. This makes it possible for us to cook tomatoes, fruits, rhubarb and use vinegar in any pan. However, fruits should not be stored for any length of time in aluminum or other metal containers. With prolonged time some aluminum or iron might dissolve and contaminate the food.

$$Fe\ (s) + 2HCl\ (aq) \longrightarrow FeCl_2\ (aq) + H_2\ (g)$$

$$2Al\ (s) + 6HCl\ (aq) \longrightarrow 2AlCl_3\ (aq) + 3H_2\ (aq)$$

Here are some examples of household acids:

Acid	**Formula**	**Common Name**	**Use**
Acetic acid	CH_3COOH		vinegar
Boric acid	H_3BO_3		eye washes
Carbonic acid	H_2CO_3		carbonated beverages, the "fizz
Citric acid	$C_6H_8O_7$		preservative
Hydrochloric acid	HCl	muriatic acid	Stomach acid; cleaning masonry
Phosphoric acid	H_3PO_4	naval jelly	provides tart taste in cola drinks; for rust removal
Potassium hydrogen tartrate	$KHC_4H_4O_6$	cream of tartar	Reacs with baking soda to make baked products rise
Sulfuric acid	H_2SO_4	battery acid	Electrolyte in automotive batteries

Bases

Bases are compounds that:

- have a bitter taste
- are neutralized by acids in a neutralization reaction in which salt and water are produced
- turn the indicator dye litmus from red to blue
- produce hydroxide ions in aqueous solutions
- have a pH above 7.0
- feel slippery or soapy on the skin

Typical bases are sodium hydroxide, NaOH, sometimes called lye; potassium hydroxide, KOH; calcium hydroxide, $Ca(OH)_2$; magnesium hydroxide, $Mg(OH)_2$; and ammonia, NH_3. All of these are solid compounds except ammonia, which is a gas. Note that bases are typically metals plus the hydroxide ion, but other compounds, such as carbonate or bicarbonate compounds) can have basic characteristics.

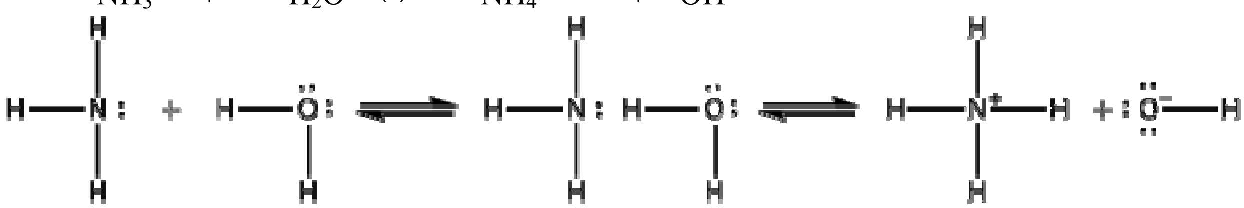
Properties of bases in water are due to the <u>hydroxide</u> <u>ion</u>, OH^-. When bases such as sodium hydroxide, potassium hydroxide or calcium hydroxide are dissolved in water, they produce hydroxide ions. They are ionic compounds and dissociate in water.

$NaOH\ (aq) \longrightarrow Na^+\ (aq) + OH^-\ (aq)$

$Mg(OH)_2\ (aq) \longrightarrow Mg^{2+}\ (aq) + 2OH^-\ (aq)$

Ammonia, NH_3, is an exception among the standard bases since it does not contain hydroxide in its formula. However, ammonia gas readily dissolves in water to produce ammonium ions, NH_4^+ and hydroxide ions, OH^-. The hydrogen leaves its electron behind when it leaves water. Therefore, the OH^- has a negative charge because it has an extra electron.

$NH_3 + H_2O \leftrightarrow NH_4^+ + OH^-$



Ammonia + Water <-----> Ammonium ion +Hydroxide Ion
Proton Acceptor

The hydroxide ion, OH^-, or a base readily <u>accepts</u> a hydrogen ion, H^+, or a proton, to produce a neutral water molecule. Bases are therefore referred to as <u>proton</u> <u>acceptors</u>.

A <u>strong</u> <u>base</u> is one that readily accepts protons. A <u>weak</u> <u>base</u> is one that has little tendency to accept protons. The classification of bases as strong or weak is a measure of the degree of reaction with hydrogen ions or protons.

Here are examples of household bases:

Base	Formula	Common Name	Uses
Ammonia	NH_3		household cleaners
Aluminum hydroxide	$A1(OH)_3$		active antacid ingredient in Maalox
Magnesium hydroxide	$Mg(OH)_2$	Milk of Magnesia	antacid
Calcium carbonate	$CaCO_3$	Limestone, calcite	antacid ingredient in Tums
Calcium hydroxide	$Ca(OH)_2$	Slaked lime	hair remover
Potassium hydrogen phosphate	K_2HPO_4		in powdered coffee creamer)
Sodium bicarbonate	$NaHCO_3$	baking soda	leavening agent
Sodium carbonate	Na_2CO_3	washing soda	laundry additive
Sodium hydroxide	NaOH	lye	oven and drain cleaners; hair relaxers

Neutralization Reactions

When an acid reacts with base, the properties of both the acid and the base are completely changed. Such reactions are referred to as neutralization reactions. The products in a neutralization reaction are a salt and water. It is a double replacement equation.

$$\underset{\text{proton donor}}{\text{Acid}} + \underset{\text{proton acceptor}}{\text{Base}} \xrightarrow{\text{Neutralization}} \text{a salt} + \text{water}$$

$$HCl + KOH \longrightarrow KCl + H_2O$$

$$H_2SO_4 + Mg(OH)_2 \longrightarrow MgSO_4 + 2H_2O$$

The acid is a compound that produces hydrogen ions, H^+, in solution and acts as a PROTON DONOR.

The base is a compound that produces hydroxide ions, OH^-, in solution and acts as a PROTON ACCEPTOR.

Consequently, the net ionic reaction between an acid and a base is:

$$\underset{\text{from acid}}{H^+} + \underset{\text{from base}}{OH^-} \longrightarrow H_2O$$

Salts (produced in the neutralization reaction), such as sodium chloride, NaC1; potassium chloride, KC1; calcium sulfate, $CASO_4$; magnesium nitrate, $Mg(NO_3)_2$ are completely dissociated in aqueous solution. A salt is an ionic compound that is generally not considered as an acid or base. Consequently, aqueous solutions of salts are good conductors of electricity.

$$NaCl\ (aq) \longrightarrow Na^+\ (aq) + Cl^-\ (aq)$$

$$Mg(NO_3)_2\ (aq) \longrightarrow Mg^{2+}\ (aq) + 2NO_3^-\ (aq)$$

Considering what we know about acids, bases and salts, we can re-write a neutralization reaction in its ionic form. The two neutralization reactions from the previous page can be written as:

Ionic equation: $H^+ (aq) + Cl^- (aq) + K^+ (aq) + OH^- (aq) \longrightarrow H_2O + K^+ (aq) + Cl^- aq)$

Net ionic equation: $H^+ (aq) + OH^- (aq) \longrightarrow H_2O$

Ionic equation: $2H^+ (aq) + SO_4^{2-} (aq) + Mg^{2+} (aq) + 2OH^- (aq) \longrightarrow$
$2H_2O + Mg^{2+} (aq) + SO_4^{2-} (aq)$

Net: $2H^+ (aq) + 2OH^- (aq) \longrightarrow 2H_2O$ or
$H^+ (aq) + OH^- (aq) \longrightarrow H_2O$

Here are some examples of neutralization reactions that occur in daily life

1. Use of antacids in the treatment of hyperacidity:

Antacids		Stomach acid		
$Mg(OH)_2$	+	$2HCl$	--->	$MgCl_2 + 2H_2O$
$Al(OH)_3$	+	$3HCl$	--->	$AlCl_3 + 3H_2O$
$CaCO_3$	+	$2HCl$	--->	$CaCl_2 + H_2O + CO_2$

2. Use of baking powder to treat battery acid spills:

$2NaHCO_3 + H_2SO_4 \longrightarrow Na_2SO_4 + 2H_2O + 2CO_2$
baking powder — battery acid — carbon dioxide

3. Effect of acid rain on limestone or marble:

$CaCO_3 + H_2SO_4 \longrightarrow CaSO_4 + H_2O + CO_2$
marble — component of acid rain

4. Effect of powdered creamer in coffee:

$K_2HPO_4 + H^+ \longrightarrow KH_2PO_4 + K^+$
base (proton acceptor) — acids in coffee

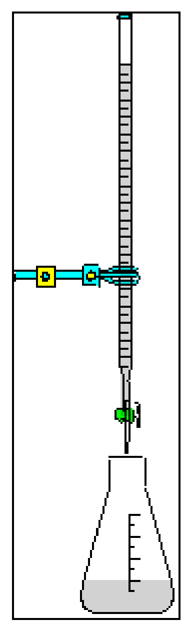
If you mixed solutions of hydrochloric acid (clear, colorless) and sodium hydroxide (clear, colorless) you would get salty water (clear, colorless). If your aim was to neutralize all the acid by adding base, how would you know that you had added enough base if there is no observation (visual or olfactory) to show that the reaction is over?

This situation occurs with many acids and bases. How do we know when an acid is neutralized by base or vice versa? We use indicators, compounds that indicate whether the solution is acidic or basic by a specific color change. There are many naturally occurring indicators. When acidic lemon juice is added to tea, it becomes lighter in color. The color change is due to the presence of the indicator in the tea. The dark color of the tea can be restored by adding a little household ammonia or some baking soda that act as a base (but you wouldn't want to drink it!). The juices of purple cabbage as well as extracts from many flowers are effective indicators. Frequently, however, the color change in any of these indicators is difficult to see. In the lab there are special color-changing dyes that are used as indicatoes. Some of these are listed below.

Indicator	Color in acid	Color in base
Methyl orange	red	yellow
Phenol red	yellow	red
Bromothymol blue	yellow	blue
Phenolphthalein	colorless	red
Litmus	red	blue

So if phenolphthalein is added to hydrochloric acid initially, the acid will remain colorless, but once enough base has been added to neutralize all the acid and one extra drop of base is added, the mixture will turn red. The appearance of the red color (still clear) shows that all the acid is gone.

Titration



Because acids and bases are so important in consumer and industrial products and processes, it is important to have an accurate means of determining the amount of acid or base in a substance. Chemists use a technique called titration. Titration is the measured addition of acid to a set amount of base until the solution is neutral or the measured addition of base to a set amount of acid until the system is neutral. A piece of glassware called a buret (burette) was designed to perform titrations. The picture to the right shows a buret suspended over an Erlenmeyer flask to do a titration. Of course, an indicator is required to let you know when the solution is neutral.

pH

The pH scale, from the French pouvoir hydrogen ("hydrogen power"), is a logarithmic scale used to express the degree of acidity or basicity (alkalinity) of a solution. The term pH is defined as a negative logarithm of the concentration of hydrogen ions or:

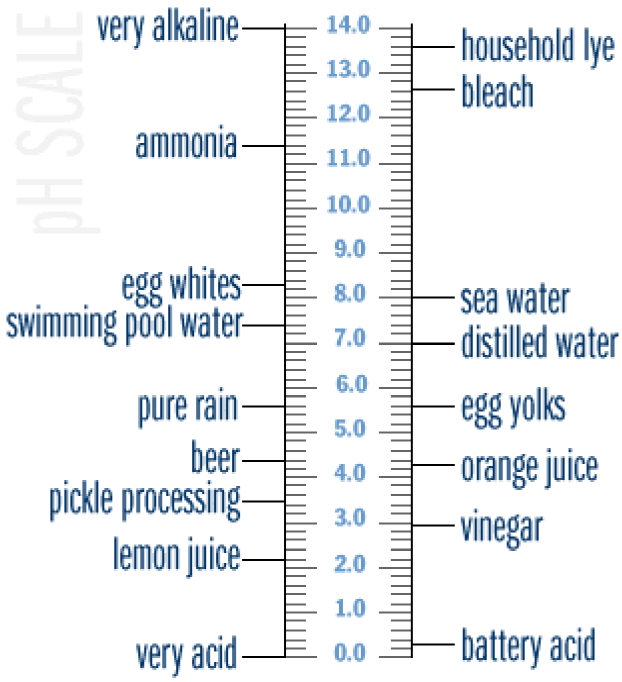
$$pH = -\log\left[H^+\right]$$

Each step on the pH scale corresponds to a tenfold change in the concentration of hydrogen ions. In other words, a pH of 2 means a hydrogen ion concentration of 10^{-2} or 0.01 mo1/L; a pH of 4 means a hydrogen ion concentration of 10^{-4} or 0.0001 mol/L. The concentration of hydrogen ions at pH = 2 (0.01) is a hundred times greater than at pH = 4 (0.0001).

pH is an easy, convenient expression commonly used to indicate the degree of acidity or basicity. A pH of 7 represents a neutral solution. A pH lower than 7 means that the solution is acidic; a pH higher than 7 indicates that the solution is basic. The lower the pH, the more acidic is the solution; the higher the pH, the more basic the solution.

Neutral

On graphic below the pH values of common solutions are given. You can see that pure distilled water and rainwater have different pH's. This is due to the acids contained in them.



↑ increasing alkalinity

↓ increasing acidity

http://www.odec.ca/projects/2005/wali5s0/public_html/pH_scale.htm

<u>Pure water</u> pH = 7

<u>Normal Rain Water</u> pH = 5.6-6.5 due to dissolved carbon dioxide

$$CO_2 + H_2O \text{ -----> } H_2CO_3 \text{ ---- > } H^+ + HCO_3^-$$

<u>Acid Rain</u> pH < 5.5 due to dissolved gases other than carbon dioxide

The record low pH for rain water occurred in a storm Scotland in 1964. The pH reading was 2.4 due to nitrogen oxides and sulfur dioxide in the atmosphere produced by combustion processes – burning coal, oil, or wood.

$$2SO_2 + O_2 + 2H_2O \longrightarrow 2H_2SO_4$$

Sulfur dioxide

Both acids make rain into "acid rain".

$$4NO + 3O_2 + 2H_2O \longrightarrow 4HNO_3$$

Nitrogen monoxide

Buffers

Most biochemical reactions in living systems are very sensitive to the pH of the solution in which they occur. Proteins can change their shapes and therefore lose their functions when the pH changes. The pH of your bloodstream must remain between 7.35 and 7.45. What happens if you drink a big glass of orange juice or lemonade? Do you go into "acidosis" (which can lead to death)?

Our body systems use solutions that contain buffers to maintain homeostatis. A buffer is a solution that resists change in pH. To do this it must have a component that can react with an acid and one that can react with a base. Sometimes one compound, such as an amino acid, can do both because it contains an acid and a base group in its structure. The acids and bases must be weak or the buffer will do more damage than good. There are many different buffer systems that operate in your body that maintain different pHs depending on the reactions in that area.

Manufacturers of certain aspirin products urge us to buy their brand because it is "buffered." Some people with a tendency toward hyperacidity may find that aspirin upsets their stomachs. So when using a buffered product, even though additional acid (aspirin is acetylsalicylic acid) is being ingested the buffer will help prevent additional acidity.

C. **Misconceptions**:

1. "All acids are dangerous." The dangerous acids are those strong acids that completely ionize in water (HC, H_2SO_4, HNO_3, and $HClO_4$ – perchloric). The others vary in strength, but most are mild enough to eat!

2. Anything containing hydrogen is an acid. The acidic hydrogen must be attached to an electronegative element like oxygen or chlorine if it is to be released. Acetic acid, CH_3COOH, has four hydrogens but three are attached to carbon and they will stay strongly bonded. Only the hydrogen attached to oxygen will ionize.

3. Strength and concentration are the same. It is common for students to use the terms "strong" and "concentrated" as synonyms because we refer to drinks as strong when we really mean concentrated. The strength of an acid (or base) comes from how much it ionizes (or

dissociates). You can have a strong acid that is dissolved in so much water that it is too dilute to cause problems. You can also have a sample of a weak acid that is so concentrated that it can eat a hole in denim jeans.

4. Buffers are always pH 7. Many people assume that if you are buffering something you are trying to make it neutral at pH= 7. You can make a buffer that will maintain pH 2 or one that will keep pH 10. Stomach buffer maintain pH 1.5-3.

D. Warm Up Exercises

1. Sweet and Sour Chicken is a popular Chinese dish. What is the "sour" and what causes it?

2. You ate something really spicy and now you have "heartburn". What can you do for this and why does it work?

3. To protect our Earth we want to be sure our water is pure. That means we need to reduce acid rain. So what is acid rain and what makes it?

4. Our bodies are really sensitive to the acid and base we eat. How come we don't have a serious problem when we drink acidic soda (like Coke, 7-Up or Mountain Dew), orange juice or lemonade?

E. Glossary

Acetic acid	CH_3COOH or $HC_2H_3O_2$ is one of the earliest known organic compounds; vinegar is a 4-5 % solution of acetic acid in water.
Acetic anhydride	$(CH_3CO)_2O$ produces acetic acid upon hydrolysis (reaction with water).
Acid rain	rain that contains acidic oxides of sulfur and nitrogen dissolved in water to produce water below pH 5.5.
Acids	a large class of chemical substances whose water solutions have one or more of the following properties; sour taste, ability to make litmus dye turn red and to cause other indicator dyes to change characteristic colors; react with bases in neutralization reaction to produce salt and water.
Acidification	to acidify means to add sufficient amount of acid to a solution until it becomes acidic; indicators are used to ascertain that the solution is acidic.
Antacids	are compounds that are used to neutralize stomach acid.

Ascorbic acid — $C_6H_8O_6$ a white solid organic acid; it is better known as vitamin C; must be present in the diet of man to prevent scurvy.

Bases — a large class of compounds with one or more of the following properties; bitter taste, slippery feeling in solution, ability to turn litmus blue and to cause other indicators to take on characteristic colors; react with acids in a neutralization reaction to produce salt and water.

Bicarbonate — a compound that contains the HCO_3^- group, i.e., sodium bicarbonate is $NaHCO_3$ and calcium bicarbonate is $Ca(HCO_3)_2$.

Buret — glassware designed for titrations; liquid is dispensed from the bottom

Buffer — solution that resists change in pH; contains components that can react with an acid and a base

Citric acid — $HOOC_2^-C(OH)^-(COOH)^-CH_2^-COOH$ is a solid organic acid; one of the most widely distributed naturally occurring acid; particularly abundant in citrus fruits; widely used in the food industry and in the preparation of beverages because of its solubility in water and mildly sour taste.

Effervescence — "bubbling," "fizzing," or appearance of gas bubbles in a solution; common occurrence when carbonate and bicarbonates react with acid

End point — the point during a titration at which a marked color change is observed, indicating that no more titrating solution is to be added.

Fatty acids — organic acids with large molar mass found in fats or lipids; general formula is R-COOH where R is a straight chain hydrocarbon portion.

Hydrochloric acid — aqueous solution of hydrogen chloride, HC; hydrochloric acid is a strong inorganic acid known as stomach acid.

Indicator — an organic substance (usually a dye or an intermediate) which indicates the presence or absence or concentration of some other substance by a change in its color; the most common example is the use of acid-base indicators such as litmus, phenolphthalein, and methyl orange to indicate the presence or absence of acids and bases.

Litmus — an indicator that appears red in acidic medium and blue in basic medium.

Neutralization	reaction of an acid with a base in which a salt and water are produced;

$$\text{Acid} + \text{base} \dashrightarrow \text{salt} + \text{water}$$
$$H_2SO_4 + Ca(OH)_2 \text{ ---> } CaSO_4 + 2H_2O$$

pH	a measure of acidity or basicity of a compound; it is the negative logarithm of the concentration of hydrogen ions; pH = -log $[H^+]$; pure water, which is neutral has pH = 7; acids have pH lower than 7, while bases have pH greater than 7; the stronger the acid, lower the pH and the stronger the base, higher the pH.
Salt	the compound that is produced when an acid reacts with a base in a neutralization reaction; for example:

$$HCl + NaOH \text{ ---> } \underset{\text{Salt}}{NaC1} + H_2O$$

Sodium hydroxide	NaOH known as caustic soda or lye; it is a strong inorganic base and readily neutralizes acid
Sulfuric acid	H_2SO_4 a strong inorganic acid that dissolves most metals; also known as battery acid; most widely used industrial chemical
Titration	slow addition of an acid to a base or vice-versa in the presence of an indicator until the end point is reached; used for quantitative analysis of acid and base solutions.

CHAPTER 9

Chemistry of Everyday Life

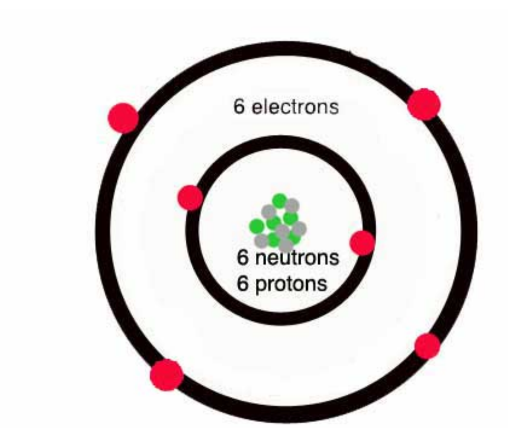
Organic Chemistry

B. Background

Many of the compounds we are in daily contact with are organic compounds, compounds which contain carbon. Organic chemistry is the study of these compounds. There are many more carbon compounds than compounds of all the other 118 elements combined. Carbon rates its own branch of chemistry because of the unique property that it can form chains and sheets of several atoms to several thousand atoms. The chains and sheets can be flexible or rigid. The chains can also form into rings or chains of rings. Carbon will also bond covalently with oxygen, hydrogen, nitrogen, sulfur and the halogens (F, Cl, Br, I) which add diversity to the array of organic compounds.

Organic Chemistry

Carbon (C) is a non-metal located in Period 2 and Group IVA on the Periodic Table.



1 http://www.historyforkids.org/scienceforkids/chemistry/atoms/pictures/carbon.jpg

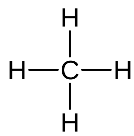
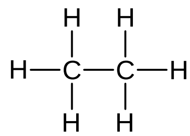
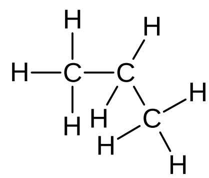
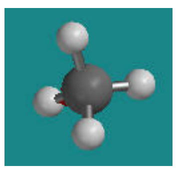
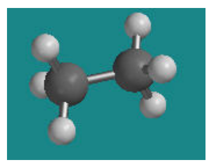
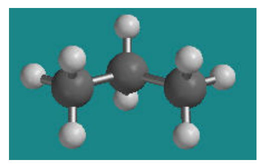
Since carbon has 4 electrons in its outer or valence shell, it could either lose 4 e^-s or gain 4 e^-s to attain a complete outer shell. Instead, it shares electrons with other atoms. Each carbon atom must share a total of 4 e^-s.

The diagrams used in this chapter to illustrate organic compounds are structural formulas in which the covalent bond or the sharing of 2 electrons between atoms is indicated by a dash. In these structural formulas, carbon must always have a total of four dashes or bonds. Hydrogen can only have one dash between itself and another atom, since it only needs two electrons to complete its shell since the shell is closest to the nucleus. Oxygen and sulfur with six valence electrons share 2 electrons and have 2 dashes. Nitrogen has 5 electrons, so it bonds by sharing 3 electrons with other atoms (3 dashes). Structural formulas not only tell us which elements are in a compound and how many atoms of each are included, but also show which atoms are bonded to which. It should be noted however, that we write these structural formulas on flat sheets of paper, but that does not mean that the molecules they represent are flat.

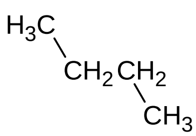
Hydrocarbons

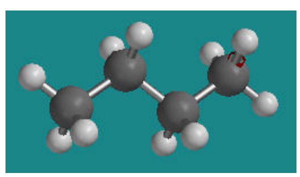
Organic compounds containing only carbon and hydrogen are referred to as hydrocarbons. They are the simplest organic compounds in composition but are not always simple in structure.

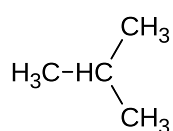
If all the carbon-carbon bonds are single bonds (sharing only one pair of electrons) the hydrocarbon is referred to as an alkane. The simplest alkane is CH_4, methane. In alkanes, the number of hydrogens is always twice the number of carbons, plus 2 more or the formula is C_nH_{2n+2}.

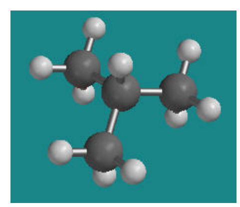
Molecular Formula	(CH_4)	(C_2H_6)	(C_3H_8)
Structural Formula			
			
Name	Methane	Ethane	Propane

Once there are 4 or more carbons in a compound, the carbons may be arranged in more than one way.





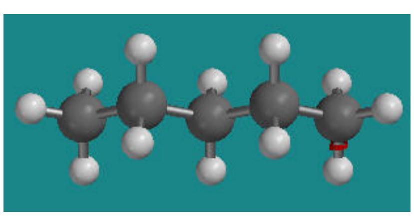


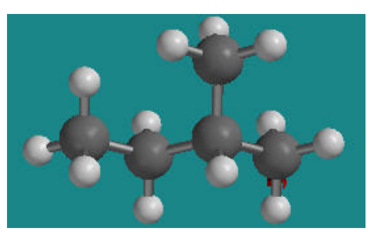


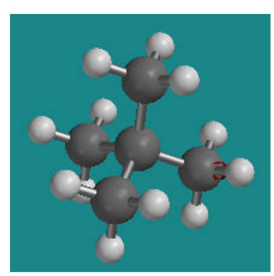
C_4H_{10}
Butane

C_4H_{10}
Isobutane

For 5 carbons (C_5H_{12}) , the following structures are possible:

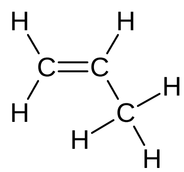




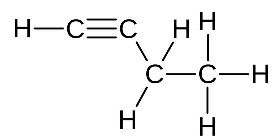


Compounds which have the same molecular formula but different structural formulas are called <u>isomers</u>. In organic chemistry structural formulas are used almost exclusively because so many isomers exist. Although the molecular formulas are the same, the differences in the shape of isomers result in differences in physical properties such as boiling and freezing points.

Notice that for all compounds written so far each carbon has formed 4 bonds to 4 other atoms. No additional atoms can be accepted by the carbon. We say that these compounds are <u>saturated</u>. Carbon can be bonded to only 3 or 2 additional atoms as seen here:



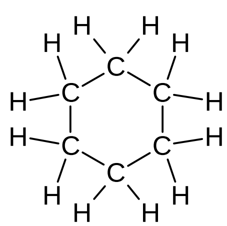
Double bond
C shares 2 pair of electrons

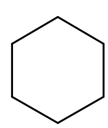


Triple bond
C shares 3 pairs of electrons

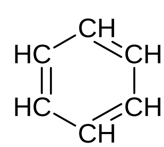
But each individual carbon still has a total of four bonds (4 dashes), as a result of sharing more than one e^- with another atom. Such compounds are referred to as <u>unsaturated</u> hydrocarbons because the double or triple bonds can be broken allowing more atoms to be added to carbon. When something is added to a double or triple bond it is referred to as an <u>addition</u> <u>reaction</u>.

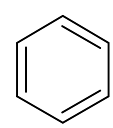
Organic compounds can form rings as well as chains.





Cyclohexane C_5H_{12}



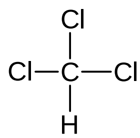


Benzene C_6H_6

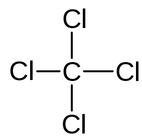
The benzene ring, containing 6 carbons bonded by alternating single and double bonds, illustrated above is a particularly stable ring and is found often in nature. The simple rings shown next to the structures are abbreviated forms of these molecules.

<u>Additions to Carbon Chains and Rings</u>

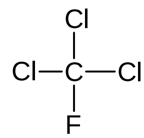
As mentioned before, organic compounds can contain more than C and H. When a halogen (F, C1, Br, I) is attached to a carbon in place of a hydrogen the compound becomes a <u>halohydrocarbon</u> <u>compound</u>. More specifically these compounds will be called fluorocarbons if F is attached, chlorocarbons if C1 is attached and chlorofluorocarbons if C1 and F are replacing hydrogens. Below are some common halohydrocarbons:



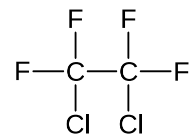
Chloroform



Carbon Tetrachloride



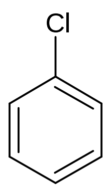
Freon 11®



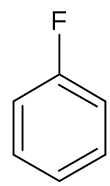
Freon 114®

The Freons® are used as refrigerants and propellants in aerosol cans. However, because they have been implicated in the destruction of the ozone layer, their use and production is being phased out.

C1 and F can also substitute for hydrogen attached to carbon rings.



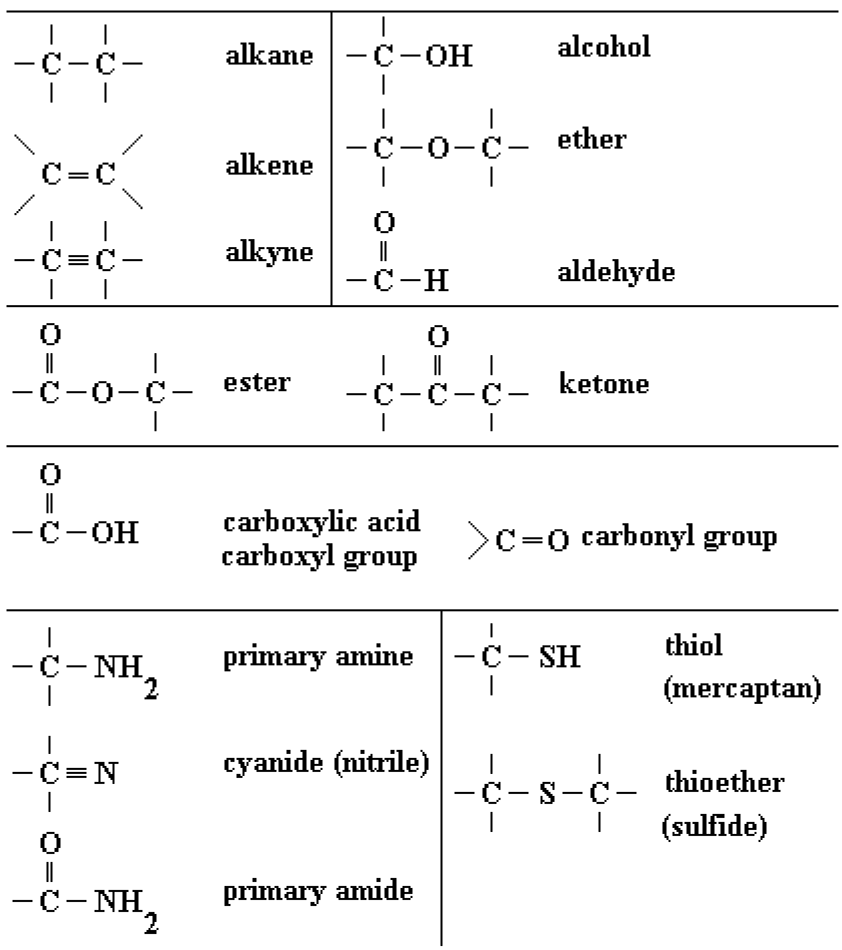
Chlorobenzene



Fluorobenzene

Functional Groups

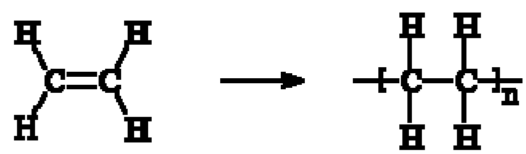
Many times a particular group of atoms or bonds will appear in an organic compound. If this group has its own unique reactions independent of how many carbons are in the chain, we refer to it as a <u>functional group</u>. A list of selected functional groups is shown to the right. Many of the tests (or unique reactions) for each group, provide the basis for analyzing different drugs or determining whether a food contains carbohydrates, lipids or proteins.



Polymers

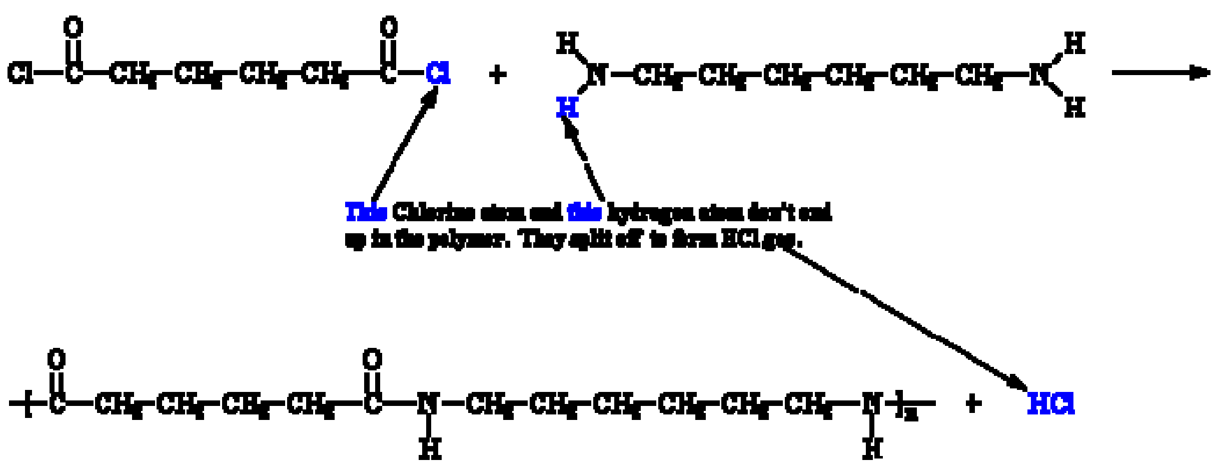
One unique property of carbon is its ability to form very large molecules or macromolecules. Examples of natural macromolecules include DNA, proteins, starches, cellulose (fiber in your diet). Man-made macromolecules include Styrofoam, polyvinyl chloride (PVC) and Orlon®. The general public has often applied the term plastic to describe man-made giant molecules. This is incorrect because in scientific terms a plastic is any substance that can be softened by heat and formed by pressure. Chemists prefer the term polymer to describe a man-made macromolecule. A polymer is a macromolecule formed from small repeating units called monomers. The nature of the polymer is very different from the monomer. The polymer can be made from all the same monomer. It can also be formed by repeating two or more different monomers, in which case it is called a copolymer. The process of forming the polymer is called polymerization.

Polymers form in two basic ways, addition and condensation. In addition polymerization double bonds in the monomer units are broken so that the monomers can join.



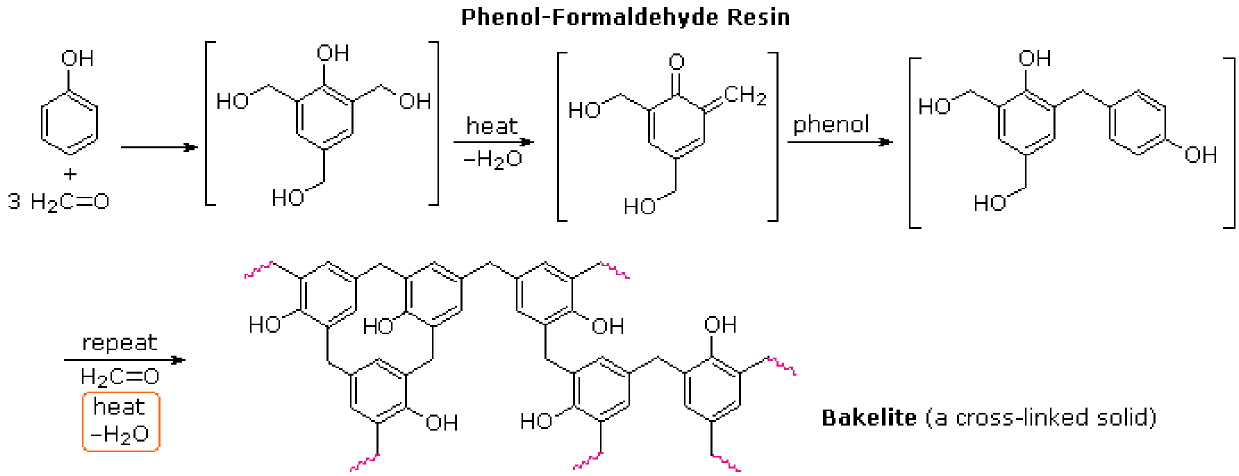
Ethylene has two carbon atoms and four hydrogen atoms, and the polyethylene repeat structure has two carbon atoms and four hydrogen atoms. None gained, none lost.

In condensation polymerization, a part of each monomer is removed and the rest of the monomer pieces are joined. This continues on both ends as the polymer builds up.



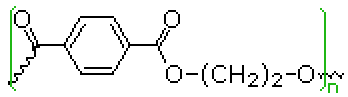
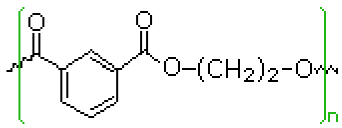
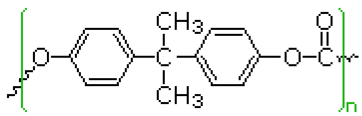
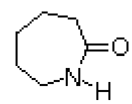
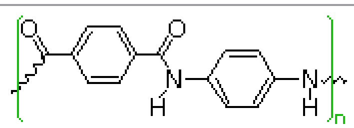
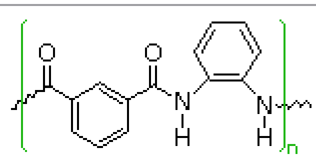
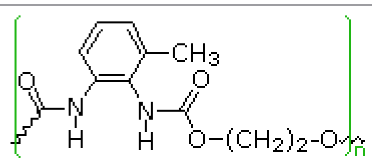
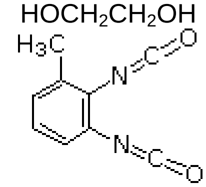
Polyesters (Dacron®, Mylar®), polyamides (Nylons) and polycarbonates (Lexon®) are examples of condensation polymers.

Polymers can also be cross-linked with bonds between chains:



Commercial polymers are designed to meet particular requirements such as rigidity or flexibility, transparency or reshaping. The monomers used and degree of cross-linking vary. Any polymer that can be heated and remolded is called a thermoplastic. A plastic that is permanently set by heat and/or pressure is called a thermosetting plastic. Additives may be included as the polymer is formed to enhance or create the properties desired.

Some Common Addition Polymers				
Name(s)	**Formula**	**Monomer**	**Properties**	**Uses**
Polyethylene low density (LDPE)	$–(CH_2\text{-}CH_2)_n–$	ethylene $CH_2{=}CH_2$	soft, waxy solid	film wrap, plastic bags
Polyethylene high density (HDPE)	$–(CH_2\text{-}CH_2)_n–$	ethylene $CH_2{=}CH_2$	rigid, translucent solid	electrical insulation bottles, toys
Polypropylene (PP) different grades	$–[CH_2\text{-}CH(CH_3)]_n–$	propylene $CH_2{=}CHCH_3$	atactic: soft, elastic solid isotactic: hard, strong solid	similar to LDPE carpet, upholstery
Poly(vinyl chloride) (PVC)	$–(CH_2\text{-}CHCl)_n–$	vinyl chloride $CH_2{=}CHCl$	strong rigid solid	pipes, siding, flooring
Poly(vinylidene chloride) (Saran A)	$–(CH_2\text{-}CCl_2)_n–$	vinylidene chloride $CH_2{=}CCl_2$	dense, high-melting solid	seat covers, films
Polystyrene (PS)	$–[CH_2\text{-}CH(C_6H_5)]_n–$	styrene $CH_2{=}CHC_6H_5$	hard, rigid, clear solid soluble in organic solvents	toys, cabinets packaging (foamed)
Polyacrylonitrile (PAN, Orlon, Acrilan)	$–(CH_2\text{-}CHCN)_n–$	acrylonitrile $CH_2{=}CHCN$	high-melting solid soluble in organic solvents	rugs, blankets clothing
Polytetrafluoroethylene (PTFE, Teflon)	$–(CF_2\text{-}CF_2)_n–$	tetrafluoroethylene $CF_2{=}CF_2$	resistant, smooth solid	non-stick surfaces electrical insulation
Poly(methyl methacrylate) (PMMA, Lucite, Plexiglas)	$–[CH_2\text{-}C(CH_3)CO_2CH_3]_n–$	methyl methacrylate $CH_2{=}C(CH_3)CO_2CH_3$	hard, transparent solid	lighting covers, signs skylights
Poly(vinyl acetate) (PVAc)	$–(CH_2\text{-}CHOCOCH_3)_n–$	vinyl acetate $CH_2{=}CHOCOCH_3$	soft, sticky solid	latex paints, adhesives
Polychloroprene (cis + trans) (Neoprene)	$–[CH_2\text{-}CH{=}CCl\text{-}CH_2]_n–$	chloroprene $CH_2{=}CH\text{-}CCl{=}CH_2$	tough, rubbery solid	synthetic rubber oil resistant

Common Condensation Polymers		
Formula	**Type**	**Components**
~[$CO(CH_2)_4CO\text{-}OCH_2CH_2O$]$_n$~	**polyester**	$HO_2C\text{-}(CH_2)_4\text{-}CO_2H$ $HO\text{-}CH_2CH_2\text{-}OH$
T 	**polyester** Dacron Mylar	para $HO_2C\text{-}C_6H_4\text{-}CO_2H$ $HO\text{-}CH_2CH_2\text{-}OH$
	**polyester**	meta $HO_2C\text{-}C_6H_4\text{-}CO_2H$ $HO\text{-}CH_2CH_2\text{-}OH$
	**polycarbonate** Lexan	$(HO\text{-}C_6H_4\text{-})_2C(CH_3)_2$ (Bisphenol A) $X_2C{=}O$ (X = OCH_3 or Cl)
~[$CO(CH_2)_4CO\text{-}NH(CH_2)_6NH$]$_n$~	**polyamide** Nylon 66	$HO_2C\text{-}(CH_2)_4\text{-}CO_2H$ $H_2N\text{-}(CH_2)_6\text{-}NH_2$
~[$CO(CH_2)_5NH$]$_n$~	**polyamide** Nylon 6 Perlon	
	**polyamide** Kevlar	para $HO_2C\text{-}C_6H_4\text{-}CO_2H$ para $H_2N\text{-}C_6H_4\text{-}NH_2$
	**polyamide** Nomex	meta $HO_2C\text{-}C_6H_4\text{-}CO_2H$ meta $H_2N\text{-}C_6H_4\text{-}NH_2$
	**polyurethane** Spandex	$HOCH_2CH_2OH$ 

C. <u>Misconceptions</u>

1. There is generalized confusion concerning the terms "natural," "synthetic," "man-made," and "artificial." Many people assume that products labeled "natural" are inherently good for you and synthetic products are automatically harmful. Many chemicals found in nature (natural, not constructed or synthesized in the laboratory) are toxic – arsenic, lead, mercury for example. Ingesting too much of any natural substance can be harmful as well – so there are toxic levels of natural substances. Chemicals that exist in nature can be constructed or synthesized in the laboratory from smaller compounds or elements. Ascorbic acid, Vitamin C, occurs naturally in citrus fruit and some vegetables. It can also be made in the lab. Your body <u>cannot</u> tell the difference. There are other chemicals however that are man-made, i.e., synthesized in the laboratory, that are not found in nature such as many of the polymers discussed in this chapter. "Artificial" is a term applied to a chemical that is synthesized in the lab and is not identical in structure to the compound found in nature but has properties similar to the natural compound such as taste or color. The effect on the human body of these synthetics must be tested before they can be OK'd for consumption.

D. <u>Questions</u>

<u>Before Lesson or Lab</u>:

1. Bring in sample of materials like a leather and a synthetic belt or drinking glasses of glass and synthetic and ask kids to decide what is natural and which is synthetic without touching them. Then have them carefully handle the items and guess again. Discuss the reasons for their classifications.

2. Have a discussion of what "organic" means to them.

<u>After the Lesson</u>:

1. Make "people polymers" to represent addition polymers. You can add some people across the chains to cross-link the various sections

2. Collect polymers from around the house or classroom and see if you can identify the polymer material and some characteristics

2. Discuss the issue of the <u>pros</u> and <u>cons</u> of polymers in our lives. Remember to consider the disposal/recycling of plastics after they have served their purpose. Have a debate!

E. Glossary

Addition reaction	chemical reaction in which atoms are added to a compound without removal of other atoms
Alkane	hydrocarbon compound with only single carbon-carbon bonds
Crosslink	create a chemical bond across two or more chemical chains
Functional group	group of atoms which undergoes a specific set of reactions
Hydrocarbon	compound containing only carbon and hydrogen
Isomer	compounds with same molecular formula but different structural formula
Macromolecule	giant molecule
Monomer	small chemical compound that is joined to form a polymer
Organic chemistry	chemistry of carbon compounds
Organic compound	compound containing carbon (except CO_1, CO_2, $CO_3^{2-,}$ HCO_3^-, $CN^{-)}$
Plastic	synthetic polymer
Polymer	macromolecule mode of repeating units (monomers)
Polymerization	process of making a polymer; if monomer C=C bonds are broken to form the polymer; it is addition; if small pieces of each monomer are removed it is condensation
Saturated (organic)	unable to form additional bonds without removal of atoms
Structural formula	chemical formula which shows arrangement of atoms
Thermoplastic	polymer that can be heated and remolded
Thermosetting plastic	polymer whose shape is set by heat; cannot be remolded
Unsaturated (organic)	able to form additional bonds without removing atoms; unsaturated compounds contain double or triple bonds

F. Some Additional Resources

Here are some pages with some interesting polymer activities and information

Polymer Activities
http://www.science-house.org/CO2/activities/polymer/index.html

Polymers in action
http://pslc.ws/macrog/kidsmac/kfloor4.htm

Polymer Demonstrations and Activities
http://www.chymist.com/polymers.html

IPSE Polymer Activities
http://www.ipse.psu.edu/activities/polymers/

CHAPTER 10

Chemistry of Everyday Life

<u>Biochemistry and Food Chemistry</u>

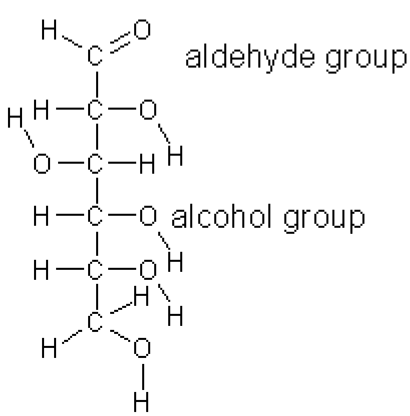
B. <u>Background</u>

<u>Biochemistry</u> is the study of the chemistry of living systems. It attempts to understand the chemical compounds and reactions that sustain life. Some of the most important advances that impact on our health and well-being have originated in this branch of chemistry. While living systems are complex, the compounds and reactions that occur in them can be understood through a sound knowledge of the processes and characteristics of organic and inorganic compounds.

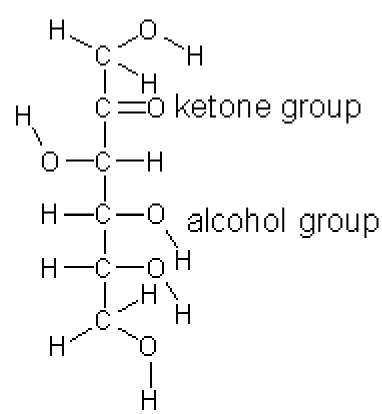
To help simplify this study of biological compounds the compounds have been classified as <u>carbohydrates</u>, <u>proteins</u>, <u>lipids</u>, and <u>nucleic</u> <u>acids</u>.

<u>Carbohydrates</u>

<u>Carbohydrates</u> contain multiple alcohol (—C—OH) groups and an aldehyde or ketone group.



Glucose
Simple sugar – (Aldose0
Monosaccharide



Fructose
simple sugar – (Ketose)
Monosaccharide

Carbohydrates can be classified as <u>monosaccharides</u> (simple sugars), <u>disaccharides</u> (2 simple sugars linked), and <u>polysaccharides</u> (many simple sugars linked). Monosacchaides cannot be decomposed easily. Disaccharides and polysaccharides can be broken into monosaccharides by <u>hydrolysis</u>, a reaction with water in which bonds are broken. Table 8.4 indicates the component parts of some common carbohydrates.

Some Carbohydrates of Dietary Importance

Carbohydrate	Hydrolyzed to	Importance
Monosaccharides		
Glucose	-	Blood sugar
Fructose	-	Fruit sugar
Galactose	-	Component of milk sugar
Disaccharides		
Maltose	Glucose	Malt sugar
Sucrose	Glucose + Fructose	Table or cane sugar
Lactose	Glucose + Galactose	Milk sugar
Polysaccharides		
Amylase	Glucose	Soluble plant starch
Amylopectin	Glucose	Insoluble plant starch
Glycogen	Glucose	Animal starch
Cellulose		Glucose fiber; plant structure material

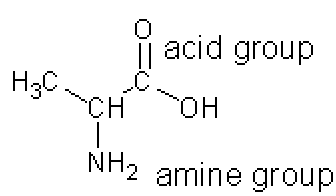
Chemical tests to distinguish the presence of carbohydrates are based on their many alcohol groups and/or the aldehyde or ketone group(s). The primary function of carbohydrates in the human system is to provide energy. This energy is made available when the carbohydrate is metabolized to CO_2 and H_2O in animal respiration.

Proteins

Proteins are macromolecules having molar masses ranging from 12,000 to 48,000 g/mol. They are vital to the body as: sources of energy; regulators of biological processes (hormones); catalysts of reactions (enzymes); transporters of oxygen (hemoglobin); defense against infection (antibodies); transmission of impulses (nerves); providers of muscular activity; buffers of the blood; components of hair, skin, and nails; as well as connective and supportive tissue.

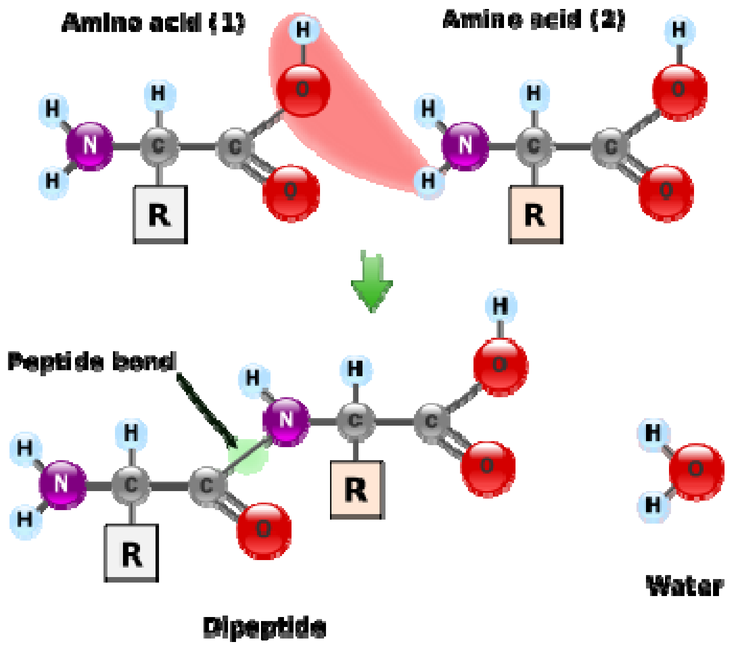
Proteins are formed from small compounds called amino acids. The general structure of an amino acid is as follows:

Example of an amino acid with acid and amine group



The body can synthesize 10 of the 20 amino acids it needs. The remaining 10 amino acids must be supplied daily by ingestion of food and are called <u>essential</u> <u>amino</u> <u>acids</u>. The tiny bacterium E. coli is far more self-sufficient than humans with respect to protein synthesis as it can make all of its own acids.

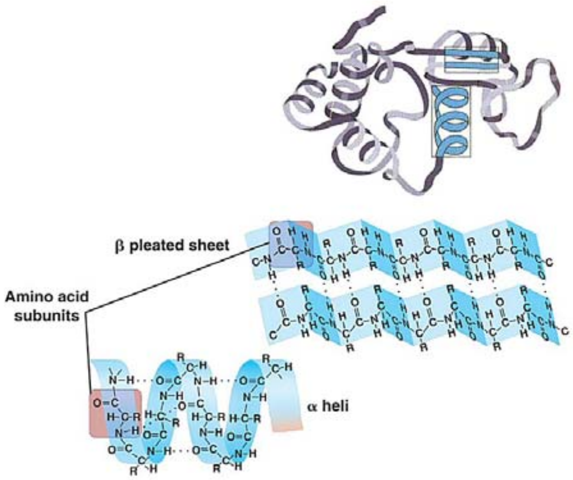
Amino acids polymerize or link together to form protein molecules. The following diagram illustrates the condensation of 2 amino acids to form the larger <u>dipeptide</u>. Note that an –OH from the acid group of one amino acid splits off with a –H from the amine portion of another amino acid and forms water. A link or bond (peptide bond) is formed between the C and N of the adjacent molecules and the larger dipeptide is formed.



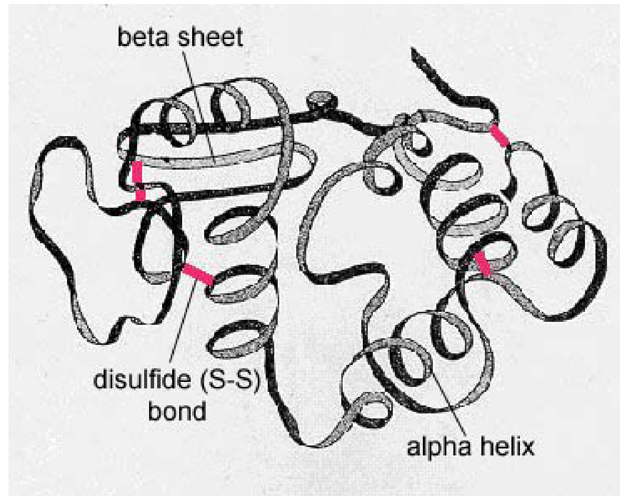
http://wpcontent.answers.com/wikipedia/commons/thumb/6/6d/Peptidformationball.svg/400px-Peptidformationball.svg.png

The combination of three amino acids is called a <u>tripeptide</u>. Proteins contain a large number of peptide linkages, and the number of possible sequences of amino acids along these chains is astronomical. The arrangements or ordering of amino acids along the chain is called the <u>primary structure</u> of a protein. Chains of amino acids can link to each other or kink or spiral due to the formation of bonds between hydrogen or sulfur atoms. These additional bonds add stability to the protein molecule and the resulting geometric pattern is termed the <u>secondary structure</u> of proteins. In some proteins, additional internal bonding causes the chain-like structure to fold in upon itself compacting, the molecule into layers or globules. This is termed the <u>tertiary structure</u>.

The final form of the large protein molecule plays an important role in its biological function. The altering of the secondary or tertiary structure of a protein is called <u>denaturing</u> and can result in the lost of biological activity. Boiling an egg denatures the albumin in the egg white.



Secondary protein structure



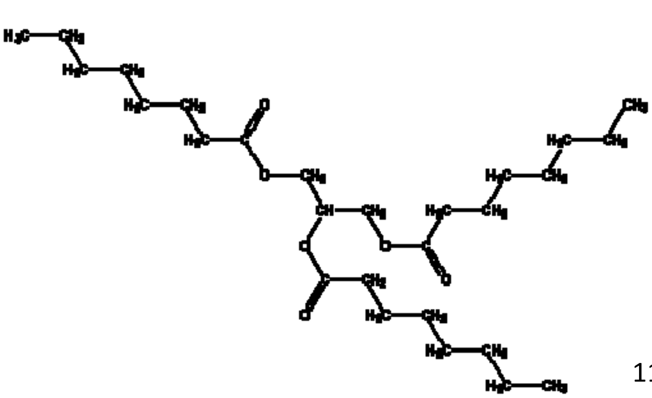
Tertiary protein structure

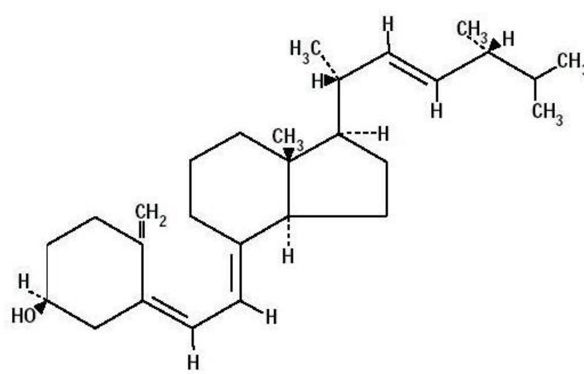
Proteins can be broken down into their component amino acids (hydrolysis) by the reaction of water, acids or bases, or enzymes.

Specific tests for the presence of proteins can be based on reactions with various amino acids or functional groups within the protein. For example, if concentrated nitric acid is added to a protein that consists of some amino acids that contain a benzene ring (tyrosine, phenylalanine, or tryptophan) a yellow color will be produced. This is the source of the yellow stains on the skin of anyone who spills nitric acid. Chromatography can be used to separate mixtures of proteins or amino acids.

Lipids

Lipid is a general term for a complex group of biochemicals that have one major characteristic in common – they are water insoluble. They are so insoluble in fact that they are termed hydrophobic (“fearing water”). Many people use the term “fat” to mean lipid. However, this is incorrect. Fats, oils, steroids, and terpenes are all classes of lipids- all are highly water insoluble. They differ in their structures and functions. In the human system lipids serve as energy storage sources, cell membrane components, hormones and emulsifiers. The tests for lipids are based on their ability to be absorbed on cellulose fibers (paper) and generate a translucent medium, and in their ability to react with iodine (if C = C groups are present in the lipid). Below are some typical lipid structures.

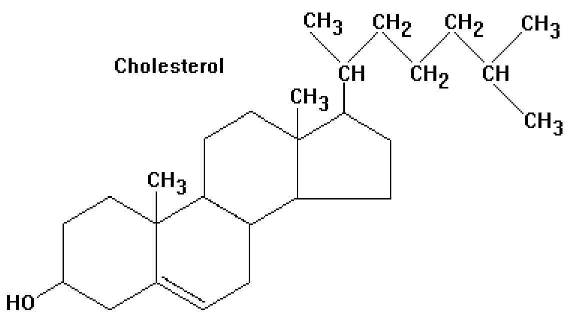




Fats and oils – are solid and liquid triacylglycerols obtained from animals and vegetables and contain mixtures of both saturated and unsaturated fatty acids.

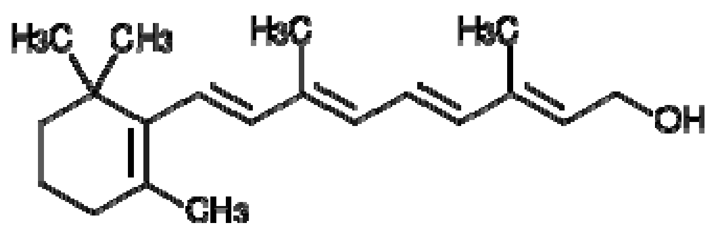
Fatty acids – such as $C_{17}H_{35}COOH$ (steric acid), are straight-chained hydrocarbons with a carboxyl group at the end. These combine with glycerol, an alcohol containing 3 –OH groups, to form triacylglyerols.

Our bodies are capable of synthesizing all but a few fatty acids from carbohydrates. Those that cannot be synthesized are called essential fatty acids – these are linoleic acid, linolenic acid, and arachidonic acid. (Sources are corn, cottonseed, peanuts, and soybean.)



Steroids are lipids that contain a characteristic carbon ring structure. Sterols, such as cholesterol, are steroids with an =OH at carbon number 3, and a branch chain of eight or more carbon atoms at carbon atoms at carbon number 17. Male and female sex hormones belong to the steroid class of compounds. The ingesting of synthetic variants of the male hormone testosterone, "anabolic steroids," has been banned in athletic competition. Such steroid increases muscle mass and endurance, but also has been implicated in damaging side effects.

Terpenes are another type of lipid consisting of characteristic isoprene units. Other terpenes include Vitamin A, E, and K.

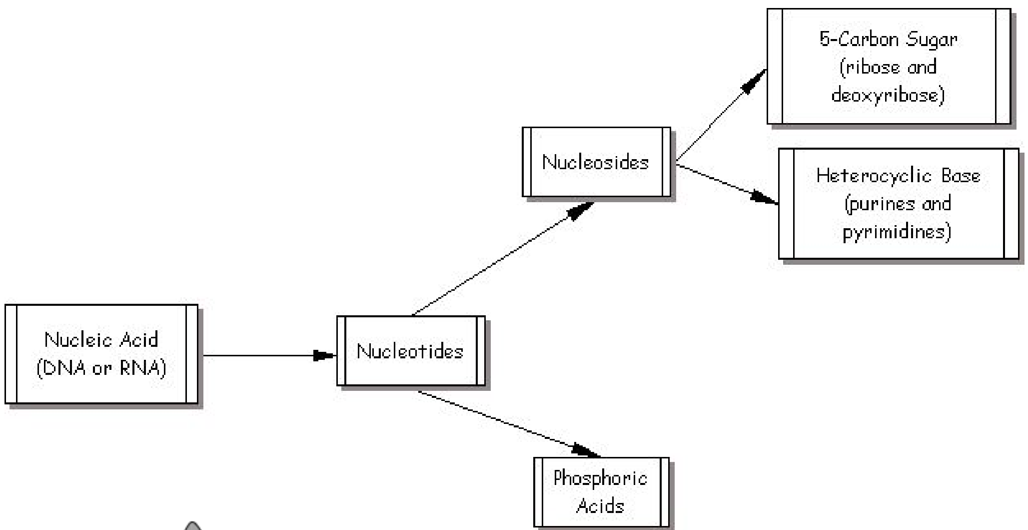


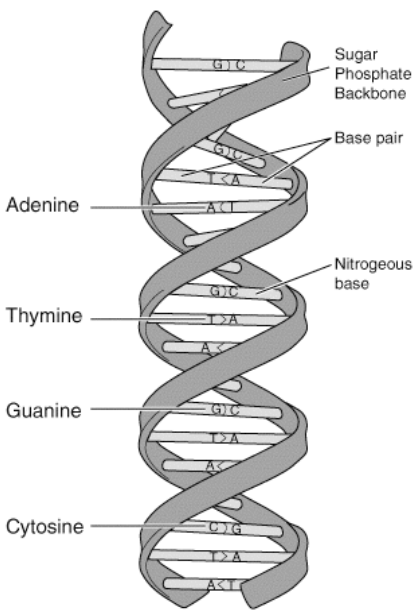
Vitamin A

Nucleic Acids

Nucleic acids are huge macromolecules – polymers with molar masses over 100 million! The units that make up the polymers are called nucleotides. Nucleotides can be broken down into nucleosides and phosphoric acid (H3PO4) – and nucleosides contain a 5-carbon sugar (ribose in RNA and deoxyribose in DNA) and a heterocyclic base (purines or pyrimidines).

DNA, the nucleic acid responsible or the transmission of genetic information, was determined by Watson and Crick to consist of 2 chains of nucleotides, with hydrogen bonding occurring between the bases on the adjacent chains. These cross-linkages between the chains cause the molecule to form into a spiral or helical shape. All DNA molecules have the same sequence of deoxyribose and phosphates in the chain, but differ in the ordering of the bases. The particular sequence of these heterocyclic bases is what constitutes the genetic code.





Section of a nucleic acid

Food Additives

Carbohydrates, proteins, and lipids are subject to chemical attack by oxidizing compounds, microbes, metals, and heat. In the processing or preparation of food, additives may be used to prevent biochemicals from undergoing chemical changes or to make the food more appealing. A food additive is a compound which has little or no nutritive value but is added to preserve or enhance the food product. Food additives may be sweeteners, coloring agents, antioxidants (compound that inhibits reactions promoted by oxygen), flavorings, and emulsifiers (promote the dispersion or mixing of liquids into liquids).

An additive cannot be used in a food product unless it meets Food and Drug Administration (FDA) approval. The agency has compiled a list of "generally regarded as safe" (GRAS) additives. Any additive that becomes suspect (carcinogenic, toxic) is removed from the list and its use prohibited.

C. Misconceptions

1. The word "food" – as is commonly used, means "stuff that plants and animals take in from the environment because they need it." This would include for humans water, and for plants, minerals from the soil and sunlight. The scientific conception of food is "organic matter that provides energy for metabolism and materials for growth."

Therefore, it is <u>incorrect</u> to say that plants obtain their food from the soil. Plants <u>make</u> all their own food. Plants use the energy from the sun and raw materials of CO_2 and H_2O to make food, but CO_2 and H_2O are not food. This confusion is compounded by the fact that plant fertilizers are labeled "plant food." When we fertilize plants, we are not <u>feeding</u> them; we are giving them raw materials so they can manufacture their own food. Humans ingest food, but H_2O is not a food.

D. Warm-up Exercises

Ingredients: Organic Corn Meal, Expeller-Pressed Sunflower Oil, Whey, Cheddar Cheese (Cultured Milk, Salt, Enzymes), Maltodextrin, Sea Salt, Natural Flavors, Disodium Phosphate, Sour Cream (Cultured Skim Milk, Cream, Cornstarch, Nonfat Dry Milk) Torula Yeast, Lactic Acid, and Citric Acid.
CONTAINS MILK INGREDIENTS

Nutrition Facts

Serving Size 1 oz.

Amount Per Serving	
Calories 150	Calories from Fat 80
	% Daily Value*
Total Fat 9g	**14%**
Saturated Fat 1.5g	**7%**
Trans Fat 0g	
Cholesterol 0mg	**0%**
Sodium 290mg	**12%**
Total Carbohydrate 16g	**5%**
Dietary Fiber less than 1g	**3%**
Sugars 1g	
Protein 2g	

Vitamin A 0% •	Vitamin C 0%
Calcium 2% •	Iron 0%
Vitamin E 10% •	Riboflavin 2%
Vitamin B_6 2% •	Phosphorus 4%

* Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs:

	Calories:	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	25g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

Calories per gram:
Fat 9 • Carbohydrate 4 • Protein 4

As students are likely to be familiar with concepts of food and nutrition, a way to prepare them for learning of the "Chemistry of Life", is to elicit their knowledge of food, food groups, and the types of nutrients. Asking questions such as "Why do we have to eat?" and "Why is it important to eat a variety of foods?" set students up to answer questions like "What are carbohydrates, proteins, and fats?"

A side panel from a box of cereal or other product can lead to a discussion of "essential" nutrients, of recommended daily requirements and the role of the Food and Drug Administration.

Additional Laboratory Exercies/Demonstrations

Some inorganic chemicals are important in living systems. Iron is a vital component of the hemoglobin protein molecule that transports oxygen.

1. <u>Fruit Juice and Tea: Testing for Iron</u>

Tea can be used to test for the presence of iron because a compound in tea forms a precipitate with iron. This causes the tea to become cloudy. It may look unpleasant but it tastes fine. This reaction can be used to test for iron in a variety of fruit juices, canned, bottled and in paper cartoons.

Procedure

1. Obtain as many test tubes or jars as juice samples and set them in a rack. Put about 1 inch (2 cm) or 5 mL of tea in each tube.

2. Add about an inch or 5 mL of one type of juice to the first test tube.

Did it get cloudy?

3. Repeat with the other juices. Record your observations in a data table. Use + for a positive (cloudy) reaction and minus for no reaction. If you are not sure if a reaction occurred, compare the pure tea and juice. Cloudy juices are harder to test.

Questions:

1. Are some juices harder to test than others? Why?
2. Are there differences among the canned, bottled, and carton juice?
3. Is iron listed as a component of the juice on its label?
4. What other things can you test besides fruit juices?

Materials and Equipment Needed:

Materials	Equipment
Fruit juices (bottled, canned, carton)	Test Tubes or colorless containers
Strong tea	

Vocabulary
Precipitate
Minimum daily requirement

Follow-up questions:

How does the body use iron?
What are the sources of dietary iron?
How much juice would you have to drink to obtain the minimum requirement?
Who decides what the MDR is? Under what circumstances might this value change?

2. Vitamin C – Food Additive

There are a number of fruits and vegetables that will turn brown on cut or bruised surfaces that are exposed to air. Apples, bananas, pears, peaches, and potatoes are examples.

The oxygen in the air reacts with a pigment in the fruit to produce the discoloration. The simplest way to prevent this is to keep the food wrapped so oxidation cannot occur. Another method is to add something to the fruit or vegetable that will react with the oxygen before the food can. Vitamin C (ascorbic acid) is a great choice.

Procedure

1. Put about 1 cup (250 mL) of water in a beaker (or bowl) and dissolve a vitamin C tablet in it.
2. Select a fruit or vegetable and quickly cut it in half. Slice half the fruit into the vitamin C solution (no core please). Make sure each piece is covered.
3. Slice the other half onto a plate exposing as much surface as possible.
4. Remove the pieces in the vitamin C solution and arrange them on a second plate.
5. Observe the slices every 5-10 minutes for the next 30 minutes.
6. Help yourself to the fruit slices!

Questions:

1. What is the difference between the treated and untreated fruit?
2. Can you think of another substance in your home that would have the same effect? (any juice with vitamin C).

Notes:

1. To prove that the water is not preventing the browning, you can set up a control and dip the second fruit half into water before putting it on the plate. However, browning takes longer this way.
2. You can experiment with rate of browning as a function of: temperature (hot vs. cold); container composition (glass or plastic vs. metal).

Materials and Equipment Needed

Materials	Equipment
Fruit	Plates
Chewable vitamin C tablet	Knife
	Slotted spoon
	Bowls or beakers

Vocabulary

Chemical change
Oxidation
Antioxidant

E. Glossary

Amino acid	compound with amine and acid functional groups; monomers for proteins
Biochemistry	chemistry of living systems
Carbohydrate	compound containing multiple -OH groups and an aldehyde or ketone group
Disaccharide	carbohydrate composed of two monosaccharides
Fatty acids	straight-chain carboxylic acids
Food	organic matter that provides energy for metabolism and materials for growth
Food additive	compound with little or no nutritive value added to preserve or enhance food product
Hydrolysis	splitting a molecule with water
Hydrophobic	highly water insoluble; water repelling
Lipid	highly water insoluble biochemical compound
Macromolecule	giant molecule
Monomer	small chemical compound that is joined to form a polymer
Monosaccharide	simple carbohydrate that cannot be decomposed by hydrolysis
Nucleic acids	macromolecules composed of nucleotide monomers that form the DNA and RNA found in the cells
Polymer	macromolecule mode of repeating units (monomers)
Polymerization	process of making a polymer; if monomer C=C bonds are broken to form the polymer; it is addition; if small pieces of each monomer are removed it is condensation
Polysaccharide	carbohydrate composed of many monosaccharides chemically linked
Protein	macromolecule made of amino acids

Steroids	a type of lipid characterized by a particular carbon ring structure
Terpenes	class of lipids consisting of isoprene units. Ex. Vitamin A
Triacylglycerols	esters of glycerol (a trihydroxy alcohol) and three fatty acids; they are lipid – and are the most common form of storage material in adipose tissue

F. Additional Resources

Lots of experiments using food with great explanations…for older students
http://extension.usu.edu/AITC/teachers/pdf/experiments_foodscience.pdf

A collection of food-based experiments
http://www.math.unl.edu/~jump/Center1/BioChemLabs.html

Suggestions for food science fair projects
http://www.juliantrubin.com/fairprojects/food/foodchemistry.html

CHAPTER 9

Chemistry At Home

B. Background

The materials used around the home, whether cleaning agents or drugs, owe their action to their chemical properties and the chemical reactions they undergo or to their physical properties. The industries that produce these products are multi-million dollar businesses that employ many chemists in the constant attempt to improve their products. An exploration of the chemistry of some of these products is discussed below.

Oxidation-Reduction Reactions

The two main classes of reactions that household products undergo are acid-base reactions and oxidation-reduction reactions. An oxidation-reduction reaction involves the transfer of electrons from one atom to another. Oxidation is loss of electrons, reduction is gain of electrons. The two processes must occur simultaneously. Below are examples of these reactions:

Overall	$CuSO_4 + Fe \text{ ------> } Cu + FeSO_4$
Oxidation	$Fe \text{ -----> } Fe^{2+} + 2\,e^-$
Reduction	$Cu^{2+} + 2\,e^- \text{ ------> } Cu$
Overall	$4\,Fe + 3\,O_2 \text{ -----> } 2\,Fe_2O_3$ (rust)
Oxidation	$4\,Fe \text{ -----> } 4\,Fe^{3+} + 12\,e^-$
Reduction	$3\,O_2 + 12\,e^- \text{ -----> } 6\,O^{2-}$

At one time oxidation was defined as combination with oxygen (as seen in the second set of reactions above). You can see from the first set that that oxygen does not have to react (or even be present) for oxidation to occur.

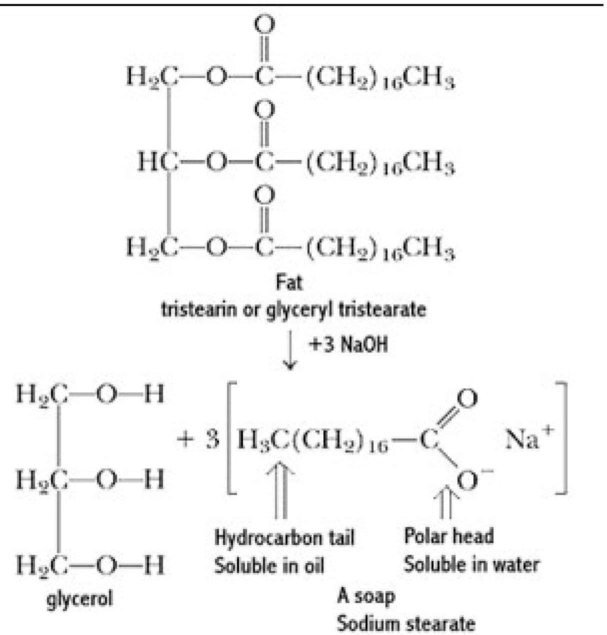
Any substance that causes oxidation is called an oxidizing agent. The oxidizing agent is reduced. A reducing agent is a substance that promotes reduction by being oxidized. In the rust reaction, oxygen is the oxidizing agent while iron is the reducing agent. The term oxidizing agent is often used with cleaning products.

Oxidation-reduction reactions, often abbreviated as redox reactions, are responsible for the rusting of iron, corrosion of metals, tarnishing and batteries (electrochemical cells).

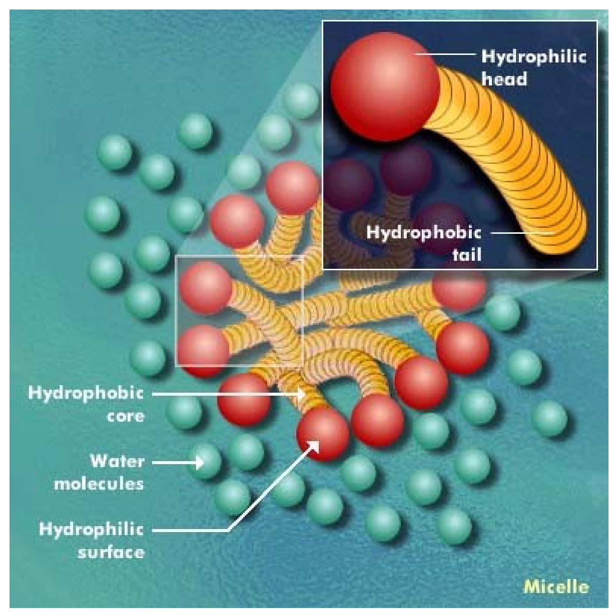
<u>Cleaning Products</u>

Cleaning products are compounds or mixtures designed to remove “dirt” or stains from a surface. Soap has been known since at least 150 A.D., while new synthetic detergents are still being produced today.

One major function of a cleaning product is to stabilize a suspension of non-polar materials such as oils or fats with a polar substance such as water. They are acting as surface-active agents, or <u>surfactants</u>. When the non-polar materials are attracted to water they can be rinsed from a surface. Soap is a surfactant composed of the sodium or potassium salts of fatty acids. It is produced when oils or fats are treated with sodium or potassium hydroxide. The process is called <u>saponification</u>:



The ionic,or <u>hydrophilic</u> end of the sodium stearate hydrogen bonds with water. The non-polar or <u>hydrophobic</u> end is repelled from water but will mix with grease. The result is grease suspended in water.



Soaps have one major drawback. They precipitate in the presence of acid or metal ions in hard water (Ca^{2+}, Mg^{2+}, Fe^{3+}). The calcium, magnesium or iron precipitates are called “scum.” For this reason, synthetic detergents were designed. These “syndets” are derived from organic products but do not produce the scum precipitate in the presence of metals. Their hydrophilic end is a polar functional group other than a carboxylic acid.

Stains can be removed from a surface by oxidizing the colored pigment to a colorless product. This is the function of bleach. The most common bleaching agent is sodium hypochlorite ($NaOCl$):

$$OCl^{-} + H_2O + 2e^{-} \text{ -----> } Cl^{-} + 2\,OH^{-}$$

The hypochlorite ion accepts electrons from the stain and is reduced while the stain is oxidized, losing its color. Hydrogen peroxide will also accomplish this. Both of these compounds however, are powerful oxidizers and will also affect textile dye. The "color-safe" bleaches contain sodium perborate which oxidizes more slowly and not affect the dye in the fabric

Fabrics can be made to appear "cleaner" by the addition of two types of compounds. Bluing agents adhere to fabrics and absorb wavelengths of light that make clothes appear yellow. This makes them appear less dingy. Optical brighteners absorb ultraviolet light and re-emit it as visible light, making clothes appear brighter. Neither of these additives has cleansing nor stain removal properties.

Cosmetic Chemistry

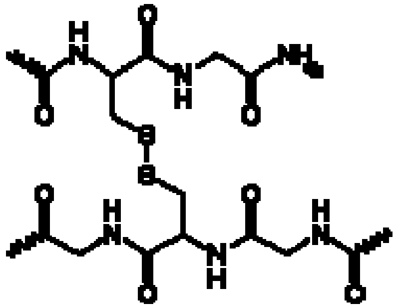
A number of household products are used to modify a person's appearance or aroma. These are classed as cosmetics and may be cleansing agents, deodorants, hair preparations, or lipsticks and powders.

Everyone has had experience with toothpaste. It is used to clean and protect tooth enamel. Tooth enamel is essentially a stone material composed of calcium hydroxy phosphate (apatite) and calcium carbonate. Both are readily attacked by acid. Acid is produced by bacteria in plaque as a by-product of sucrose or dextrin decomposition. Most tooth pastes contain an abrasive to cut surface deposits and a detergent to carry the materials away. Examples of abrasives are hydrated silica ($SiO_2 \cdot H_2O$) and calcium carbonate ($CaCO_3$). Since calcium carbonate is also a base, it can serve to neutralize acid produced in the plaque.

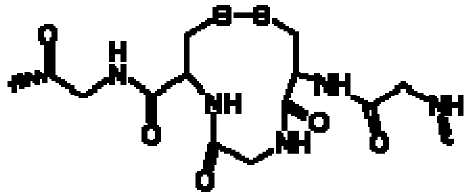
Deodorants are preparations designed to remove or mask body odor and reduce perspiration. Body odor comes from amines (proteins by-products) and fatty acids excreted from sweat glands. The odor can be reduced by using an astringent that closes the pores such as aluminum chlorohydrate ($A1_2(OH)_5C1 \cdot H_2O$). Or, you can use a product that will chemically react with the amines and fatty acids. A final approach is to use a perfume to cover the odor.

Hair products which change the color or curl in hair utilize chemical reactions to affect the changes. Hair contains two pigments, brown-black melanin and a red iron pigment. Dyes penetrate the hair fiber and enhance one or the other pigment. Bleaches oxidize the two pigments to reduce the color in the hair. Hydrogen peroxide is the most common hair oxidizing agent.

Hair curling or straightening agents work by reducing the disulfide linkages (-C-S-S-C-) in hair protein to thiols (-C-SH) and HS-C-). The hair protein chains are shifted with respect to each other and the thiols are oxidized back to disulfide groups. The bonds have now been reformed to cause curl or straightness.



disulfide crosslink



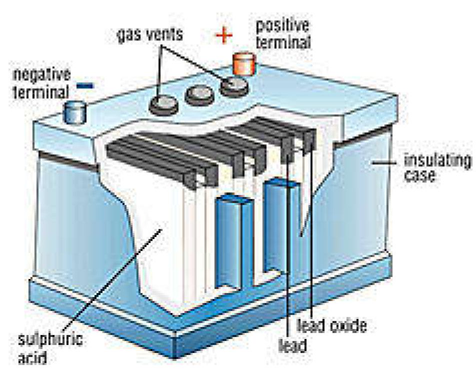
disulfide "loop"

Batteries

Oxidation-reduction reactions are the driving force behind the operation of a battery. A device that produces an electron flow (current) by means of a chemical reaction is called an electrochemical cell. A series of these cells is called a battery. (The term battery is sometimes also used when only one cell is employed.) Electrons given up by one atom at the anode, which flow to the cathode where they are accepted by a different atom. The direction in which electrons flow, that is, which atom donates and which receives, is determined by a property called electrochemical potential. Batteries "store" energy that can be used later.

Batteries in which the stored energy is used up are called primary batteries. The oxidation and reduction products are allowed to mix. Examples of primary batteries include the dry cell and alkaline dry cell. Batteries that can be recharged are called secondary batteries and the reaction products remain near their own electrodes. These can be recharged many times before losing their ability to produce a current. The most common secondary battery is the lead storage battery used in moving vehicles.

iPod battery



Lead storage battery

C. **Misconception**s:

1. Chemicals are things you use around the house or in the lab only.

All matter is chemical. The media and general public often use the term "chemical" to describe a product that has some cleaning or drug function. The connotation is often that the "chemical" is harmful.

D. **Warm-Up Exercises**

This unit may be an appropriate place to find out if students are harboring "negative connotations" derived from the media in connection with the word "chemical." Since it has been found to be very difficult to alter such negative attitudes, it may be useful to:

1. Survey the attitudes of students by using a simple Leikert-type questionnaire. One Question may be: "When I hear the word "chemical," I think of harmful "materials."

```
      5     4     3     2     1
Strongly----------------------------- > Strongly
Agree                                   Disagree
```

2. Tabulate the class results, and then ask students how they came to their decision.

3. Elicit from students a list of beneficial chemicals.

4 Discuss why many people have only the negative view. Make a distinction between chemicals, and the use of chemicals. Are there chemicals that can be beneficial and harmful depending on their use (like pain-killers)? Write a short descriptive summary of a typical morning without modern chemical products!

5. Retake the survey. Ask students to explain why they have changed their minds.

E. Glossary

Battery	series of electrochemical cells
Detergent	surface-active agent that is not made from a fatty acid
Electrochemical cell	device in which current flows; electrons are transferred in a chemical reaction
Hydrophilic	water-loving; water soluble
Hydrophobic	water repelling
Oxidation	process in which electrons are lost
Oxidation-reduction reaction	chemical reaction in which electrons are transferred
Oxidizing agent	substance that promotes oxidation; is reduced in the process
Reducing agent	substance that promotes reduction; is oxidized in the process
Saponification	alkaline hydrolysis of at or oil to produce soap
Soap	sodium or potassium salt of fatty acid
Surface tension	property of liquids in which they ‘appear’ to have an invisible coating on their surface
Surfactant	substance that reduces surface tension

Research Matters - to the Science Teacher
No. 9004 March 1, 1990

The Science Process Skills
by Michael J. Padilla, Professor of Science Education, University of Georgia, Athens, GA

Introduction

One of the most important and pervasive goals of schooling is to teach students to think. All school subjects should share in accomplishing this overall goal. Science contributes its unique skills, with its emphasis on hypothesizing, manipulating the physical world and reasoning from data.

The scientific method, scientific thinking and critical thinking have been terms used at various times to describe these science skills. Today the term "science process skills" is commonly used. Popularized by the curriculum project, Science - A Process Approach (SAPA), these skills are defined as a set of broadly transferable abilities, appropriate to many science disciplines and reflective of the behavior of scientists. SAPA grouped process skills into two types-basic and integrated. The basic (simpler) process skills provide a foundation for learning the integrated (more complex) skills. These skills are listed and described below.

<u>Basic</u> Science Process Skills

Observing - using the senses to gather information about an object or event. Example: Describing a pencil as yellow.

Inferring - making an "educated guess" about an object or event based on previously gathered data or information. Example: Saying that the person who used a pencil made a lot of mistakes because the eraser was well worn.

Measuring - using both standard and nonstandard measures or estimates to describe the dimensions of an object or event. Example: Using a meter stick to measure the length of a table in centimeters.

Communicating - using words or graphic symbols to describe an action, object or event. Example: Describing the change in height of a plant over time in writing or through a graph.

Classifying - grouping or ordering objects or events into categories based on properties or criteria. Example: Placing all rocks having certain grain size or hardness into one group.

Predicting - stating the outcome of a future event based on a pattern of evidence. Example: Predicting the height of a plant in two weeks time based on a graph of its growth during the previous four weeks.

<u>Integrated</u> Science Process Skills

Controlling variables - being able to identify variables that can affect an experimental outcome, keeping most constant while manipulating only the independent variable. Example: Realizing through past experiences that amount of light and water need to be controlled when testing to see how the addition of organic matter affects the growth of beans.

Defining operationally - stating how to measure a variable in an experiment. Example: Stating that bean growth will be measured in centimeters per week.

Formulating hypotheses - stating the expected outcome of an experiment. Example: The greater the amount of organic matter added to the soil, the greater the bean growth.

Interpreting data - organizing data and drawing conclusions from it. Example: Recording data from the experiment on bean growth in a data table and forming a conclusion which relates trends in the data to variables.

Experimenting - being able to conduct an experiment, including asking an appropriate question, stating a hypothesis, identifying and controlling variables, operationally defining those variables, designing a "fair" experiment, conducting the experiment, and interpreting the results of the experiment. Example: The entire process of conducting the experiment on the affect of organic matter on the growth of bean plants.

Formulating models - creating a mental or physical model of a process or event. Examples: The model of how the processes of evaporation and condensation interrelate in the water cycle.

Learning basic process skills

Numerous research projects have focused on the teaching and acquisition of basic process skills. For example, Padilla, Cronin, and Twiest (1985) surveyed the basic process skills of 700 middle school students with no special process skill training. They found that only 10% of the students scored above 90% correct, even at the eighth grade level. Several researchers have found that teaching increases levels of skill performance. Thiel and George (1976) investigated predicting among third and fifth graders, and Tomera (1974) observing among seventh graders. From these studies it can be concluded that basic skills can be taught and that when learned, readily transferred to new situations (Tomera, 1974). Teaching strategies which proved effective were: (1) applying a set of specific clues for predicting, (2) using activities and pencil and paper simulations to teach graphing, and (3) using a combination of explaining, practice with objects, discussions and feedback with observing. In other words-just what research and theory has always defined as good teaching.

Other studies evaluated the effect of NSF-funded science curricula on how well they taught basic process skills. Studies focusing on the Science Curriculum Improvement Study (SCIS) and SAPA indicate that elementary school students, if taught process skills abilities, not only learn to use those processes, but also retain them for future use. Researchers, after comparing SAPA students to those experiencing a more traditional science program, concluded that the success of SAPA lies in the area of improving process oriented skills (Wideen, 1975; McGlathery, 1970). Thus it seems reasonable to conclude that students learn the basic skills better if they are considered an important object of instruction and if proven teaching methods are used.

Learning integrated process skills

Several studies have investigated the learning of integrated science process skills. Allen (1973) found that third graders can identify variables if the context is simple enough. Both Quinn and George (1975) and Wright (1981) found that students can be taught to formulate hypotheses and that this ability is retained over time.

Others have tried to teach all of the skills involved in conducting an experiment. Padilla, Okey and Garrard (1984) systematically integrated experimenting lessons into a middle school science curriculum. One group of students was taught a two week introductory unit on experimenting which focused on manipulative activities. A second group was taught the experimenting unit, but also experienced one additional process skill activity per week for a period of fourteen weeks. Those having the extended treatment outscored those experiencing the two week unit. These results indicate that the more complex process skills cannot be learned via a two week unit in which science content is typically taught. Rather, experimenting abilities need to be practiced over a period of time.

Further study of experimenting abilities shows that they are closely related to the formal thinking abilities described by Piaget. A correlation of +.73 between the two sets of abilities was found in one study (Padilla, Okey and Dillashaw, 1983). In fact, one of the ways that Piaget decided whether someone was formal or concrete was to ask that person to design an experiment to solve a problem. We also know that most early adolescents and many young adults have not yet reached their full formal reasoning capacity (Chiapetta, 1976). One study found only 17% of seventh graders and 34% of twelfth graders fully formal (Renner, Grant, and Sutherland, 1978).

What have we learned about teaching integrated science processes? We cannot expect students to excel at skills they have not experienced or been allowed to practice. Teachers cannot expect mastery of experimenting skills after only a few practice sessions. Instead students need multiple opportunities to work with these skills in different content areas and contexts. Teachers need to be patient with those having difficulties, since there is a need to have developed formal thinking patterns to successfully "experiment."

Summary and Conclusions

A reasonable portion of the science curriculum should emphasize science process skills according to the National Science Teachers Association. In general, the research literature indicates that when science process skills are a specific planned outcome of a science program, those skills can be learned by students. This was true with the SAPA and SCIS and other process skill studies cited in this review as well as with many other studies not cited.

Teachers need to select curricula which emphasize science process skills. In addition they need to capitalize on opportunities in the activities normally done in the classroom. While not an easy solution to implement, it remains the best available at this time because of the lack of emphasis of process skills in most commercial materials.

Appendix A

References

Allen, L. (1973). An examination of the ability of third grade children from the Science Curriculum Improvement Study to identify experimental variables and to recognize change. *Science Education, 57*, 123-151.

Chiapetta, E. (1976). A review of Piagetian studies relevant to science instruction at the secondary and college level. *Science Education, 60*, 253-261.

McGlathery, G. (1970). An assessment of science achievement of five and six-year-old students of contrasting socio-economic background. *Research and Curriculum Development in Science Education, 7023*, 76-83.

McKenzie, D., & Padilla, M. (1984). Effect of laboratory activities and written simulations on the acquisition of graphing skills by eighth grade students. Paper presented at the annual meeting of the National Association for Research in Science Teaching, New Orleans.

Padilla, M., Okey, J., & Dillashaw, F. (1983). The relationship between science process skills and formal thinking abilities. *Journal of Research in Science Teaching, 20*.

Padilla, M., Cronin, L., & Twiest, M. (1985). The development and validation of the test of basic process skills. Paper presented at the annual meeting of the National Association for Research in Science Teaching, French Lick, IN.

Quinn, M., & George, K. D. (1975). Teaching hypothesis formation. *Science Education, 59*, 289-296. *Science Education, 62*, 215-221.

Thiel, R., & George, D. K. (1976). Some factors affecting the use of the science process skill of prediction by elementary school children. *Journal of Research in Science Teaching, 13*, 155-166.

Tomera, A. (1974). Transfer and retention of transfer of the science processes of observation and comparison in junior high school students. *Science Education, 58*, 195-203.

Wideen, M. (1975). Comparison of student outcomes for Science - A Process Approach and traditional science teaching for third, fourth, fifth, and sixth grade classes: A product evaluation. *Journal of Research in Science Teaching, 12*, 31-39.

Wright, E. (1981). The long-term effects of intensive instruction on the open exploration behavior of ninth grade students. *Journal of Research in Science Teaching, 18.*

***Research Matters - to the Science Teacher* is a publication of the National Association for Research in Science Teaching**

Appendix B

http://www.ed.gov/databases/ERIC_Digests/ed282776.html

ERIC Identifier: ED282776
Publication Date: 1987-00-00
Author: Blosser, Patricia E.
Source: ERIC Clearinghouse for Science Mathematics and Environmental Education Columbus OH.

Science Misconceptions Research and Some Implications for the Teaching of Science to Elementary School Students. ERIC/SMEAC Science Education Digest No. 1, 1987.

TEXT: INTRODUCTION

In July, 1983, an international seminar on misconceptions in science and mathematics was held at Cornell University (Helm and Novak, 1983). Fifty-five papers were presented and 118 people registered for the seminar. The proceedings of this conference were published, with the papers grouped according to primary emphasis: theoretical and philosophical perspectives (8 papers), instructional issues (9 papers), research and methodological issues (12 papers), historical and epistemological perspectives (5 papers), elementary school science (2 papers), physics (11 papers), biology (6 papers), chemistry (1 paper), and mathematics (5 papers). A second international seminar is scheduled for the summer of 1987, also at Cornell.

Although elementary school science as a primary paper emphasis accounted for only two papers, the area of misconceptions research has relevance for the teaching of science to elementary school students. This digest has been produced to describe what this area of research encompasses, to highlight a few relevant studies, and to communicate some of the implications that the findings of misconceptions research has for the teaching of science in the elementary school.

A VARIETY OF TERMS

An article published in SCIENCE EDUCATION in April 1940 was entitled "An Evaluation of Certain Popular Science Misconceptions" (Hancock, 1940). This author defined a "misconception" as "...any unfounded belief that does not embody the element of fear, good luck, faith, or supernatural intervention" (p. 208). Hancock considered misconceptions to arise from faulty reasoning. Current science education researchers would probably take issue with this assumption.

Science educators, in the United States and abroad, who are interested in conceptual development have used a variety of terms to describe the situation in which students' ideas differ from those of scientists about a concept. Some talk of students' misconceptions; others write of preconceptions; still others, of naive conceptions; some, of naive theories; some, of alternative conceptions; and some, of alternative frameworks.

Barrass (1984) wrote of "mistakes" or errors, "misconceptions" or misleading ideas, and "misunderstandings" or misinterpretations of facts, saying that teachers and brighter students can correct errors. But what attention is paid to misconceptions and misunderstandings that are perpetuated by teachers and textbook authors?

Driver and Easley (1978) contend that semantics indicate the writer's philosophical position, saying that Ausubel talks of "preconceptions," which are ideas expressed that do not have the status of generalized understandings that are characteristic of conceptual knowledge. However, those who use the term "misconception" indicate an obvious connotation of a wrong idea or an incorrectly assimilated formal model or theory. And, those persons who use "alternative frameworks" indicate that pupils have developed autonomous frameworks for conceptualizing their experience of the physical world.

Helm and Novak, in the introduction to the proceedings of the 1983 seminar, stated that an issue which surfaced early in the meeting was that "misconceptions" as a term carried with it some connotations that are not appropriate (1983). This issue was not resolved, although Novak suggested that researchers adopt the acronym LIPH, standing for "Limited or Inappropriate Propositional Hierarchies." However, seminar participants decided that it was too early in the history of research programs to attach an explicit label.

FINDINGS RELATED TO ELEMENTARY SCIENCE

What does all this mean in terms of teaching science in elementary schools? Frequently, when science is taught to elementary school pupils, it is taught as if the children had had no prior experiences relative to the topic being studied. Misconceptions research contains findings indicating that this is not a valid assumption. Children come to school already holding beliefs about how things happen, and have expectations--based on past experiences--which enable them to predict future events. They also possess clear meanings for words which are used both in everyday language and in a more specialized way in science. A child's view and understanding of word meanings are incorporated into conceptual structures which provide a sensible and coherent understanding of the world from the child's point of view (Osborne and Gilbert, 1980). Children hold ideas that were developed before and during their early school years, and these ideas may be compounded by the teacher and/or the textbook. It is possible that children develop parallel but mutually inconsistent explanations of scientific concepts--one for use in school and one for use in the "real world" (Trowbridge and Mintzes, 1985).

Fisher contends that misconceptions serve the needs of the persons who hold them and that erroneous ideas may come from strong word association, confusion, conflict, or lack of knowledge (1985). According to Fisher, some alternative conceptions, judged to be erroneous ideas or misconceptions, have these characteristics in common:

1. They are at variance with conceptions held by experts in the field. 2. A single misconception, or a small number of misconceptions, tend to be pervasive (shared by many different individuals). 3. Many misconceptions are highly resistant to change or alteration, at least by traditional teaching methods. 4. Misconceptions sometimes involve alternative belief systems

comprised of logically linked sets of propositions that are used by students in systematic ways. 5. Some misconceptions have historical precedence: that is, some erroneous ideas put forth by students today mirror ideas espoused by early leaders in the field. 6. Misconceptions may arise as the result of: a) the neurological "hardware" or genetic programming (as in the case of automatic language-processing structures, which may be invoked when "reading" an equation); b) certain experiences that are commonly shared by many individuals (as with moving objects); or c) instruction in school or other settings (p. 53).

Several reports have been produced as a result of a project carried out at the Institute for Research on Teaching at Michigan State University (Roth, 1985; Smith and Anderson, 1984a; Smith and Anderson, 1984b; Smith, 1983). This representative (but not exhaustive) list relates to using activities from the Science Curriculum Improvement Study (SCIS) with elementary school pupils. SCIS activities were not sufficient to help students exchange their previous conceptions so curriculum materials, a text, and a teacher's guide were developed for use in the project. Even when these specially developed instructional materials were used, misconceptions held by children proved difficult to change, although the modified materials were more effective than SCIS (Roth, 1985).

Operating on the assumption that, if science in the schools is to improve, elementary school science teaching has to improve, Lawrenz (1986) investigated inservice elementary school teachers' understanding of some elementary physical science concepts. She devloped a questionnaire using items from the physical science test questions given to 17-year-old students as part of the National Assessment of Educational Progress science studies, and found that 11 of the 31 items were answered correctly by 50 percent or fewer of the 333 teachers surveyed. Lawrenz concluded that some of the errors were due to lack of content knowledge, but that others were indicative of serious misconceptions. If teachers do not understand elementary physical science concepts, how can they teach their students?

IMPLICATIONS FOR TEACHING, TEACHER EDUCATION

Lawrenz (1986) advocated inservice education, beginning with very basic science concepts so that inservice teachers could have experiences with concrete examples that conflict with misconceptions they hold. Then, teachers should be shown and given numerous examples of how to identify misconceptions held by pupils in their own classrooms.

Smith and Anderson (1984b) suggested that, in teacher education programs, preservice teachers should be helped to develop ideas about conceptual change in learning. Teacher educators must realize that their students have conceptions about teaching and learning that are different from those the teacher educators hold--and that the teacher educators should work to change these students' misconceptions. They wrote:

Among the important learning outcomes teacher education should address are the following: 1. a conceptual change view of learning, 2. knowledge of generic strategies useful in achieving conceptual change, 3. knowledge of common misconceptions for several important topics and specific strategies for changing them, 4. skill in selecting and adapting curriculum materials based on common preconceptions held by students, 5. skill in diagnosing student conceptions and recognizing them from student responses, and 6. a view of theory as invented to account for observations rather than deriving objectively and reliably from them (p. 697).

Engel Clough and Wood-Robinson (1985) have suggested several things teachers may try, although they admit that these ideas have not been tested: (1) start with students' ideas and devise teaching strategies to take some account of them; (2) provide more structured opportunities for students to talk through ideas at length, both in small group and whole class discussions; (3) begin with known and familiar examples; (4) introduce some science topics into the curriculum at earlier grade levels, drawing on out-of-school knowledge (p. 129).

Several researchers have emphasized the importance of allowing pupils to explore their own ideas in a non-threatening atmosphere. Teachers need to devise strategies for encouraging this exploration and for creating the necessary classroom climate.

Teachers also need to consider the extent to which misconceptions may be language difficulties. Teachers and students may fail to share the meaning of the terms they use or the questions they ask.

Hopps (l985), in discussing cognitive learning theory and classroom complexity, has provided some suggestions that are relevant to structuring elementary school science lessons to deal with misconceptions:

--We cannot expect learners to identify and select key stimuli without specific advice from teachers
--We cannot expect that all pupils will focus attention on key aspects of the learning activity without deliberate action on the teacher's part
--Models of conceptual change imply that the learner's ability to reforge links between prior knowledge and sensory input is likely to be of critical importance in learning
--Teachers can assist learners by providing the kinds of information and experiences which will enable them to bridge the gaps between sensory input and prior knowledge...ideas to be taught should always be related to the relevant frameworks held by the learner and revision of the key parts of such frameworks should not be undertaken lightly.
--Explanations of any links between new information and prior knowledge should be made in a variety of ways so that learners are presented with visual, verbal and/or a diagrammatic format of the principles to be taught.
--Whenever concepts or definitions are to be introduced, teachers should provide significant numbers of examples and non-examples pp. 171-172).

FOR MORE INFORMATION

Barrass, Robert. "Some Misconceptions and Misunderstandings Perpetuated by Teachers and Textbooks of Biology." JOURNAL OF BIOLOGY EDUCATION 18 (1984): 201-205.

Driver, Rosalind, and Jack Easley. "Pupils and Paradigms: A Review of Literature Related to Concept Development in Adolescent Science Students." STUDIES IN SCIENCE EDUCATION 5 (l978): 61-84.

Appendix B

Engel Clough, Elizabeth, and Colin Wood-Robinson. "How Secondary Students Interpret Instances of Biological Adaptation." JOURNAL OF BIOLOGY EDUCATION 19 (1985): 125-130.

Fisher, Kathleen. "A Misconception in Biology: Amino Acids and Translation." JOURNAL OF RESEARCH IN SCIENCE TEACHING 22 (1985): 53-62.

Hancock, Cyril H. "An Evaluation of Certain Popular Science Misconceptions." SCIENCE EDUCATION 24 (1940): 208-213.

Helm, Hugh, and Joseph D. Novak. PROCEEDINGS OF THE INTERNATIOAL SEMIAR ON MISCONCEPTIONS IN SCIENCE AND MATHEMATICS. Ithaca, NY: Cornell University, July, 1983. ED 242 553.

Hopp, John C. "Cognitive Learning Theory and Classroom Complexity." RESEARCH IN SCIENCE AND TECHNOLOGICAL EDUCATION 3 (1985): 159-174.

Lawrenz, Frances. "Misconceptions of Physical Science Concepts Among Elementary School Teachers." SCHOOL SCIENCE AND MATHEMATICS 86 (1986): 654-660.

Osborne, Roger J., and John K. Gilbert. "A Technique for Exploring Students' Views of the World." PHYSICS EDUCATION 15 (1980): 376-379.

Roth, Kathleen. "Food for Plants: Teacher's Guide. Research Series No. 153." East Lansing, MI: Michigan State University, Institute for Research on Teaching, January, 1985. ED 256 624.

Smith, Edward L. "Teaching for Conceptual Change: Some Ways of Going Wrong." Final Report. East Lansing, MI: Michigan State University, Institute for Research on Teaching, June 1983. ED 237 493.

Smith, Edward L., and Charles W. Anderson. "The Planning and Teaching Intermediate Science Study: Final Report. Research Series No. 147." East Lansing, MI: Michigan State University, Institute of Research on Teaching, June, 1984a. ED 250 161.

Smith, Edward L., and Charles W. Anderson. "Plants as Producers: A Case Study of Elementary Science Teaching." JOURNAL OF RESEARCH IN SCIENCE TEACHING 21 (1984b): 685-698.

Trowbridge, John E., and Joel L. Mintzes. "Students' Alternative Conceptions of Animals and Animal Classification." SCHOOL SCIENCE AND MATHEMATICS 85 (1985): 304-316.
.

Appendix B

Some Chemistry Misconceptions

Electrical Nature of Matter

- Positively charged objects have gained protons, rather than being deficient in electrons.
- Electrons which are lost by an object are really lost (no conservation of charge).
- All atoms are charged.
- A charged object can only attract other charged objects.
- The electrostatic force between two charged objects is independent of the distance between them.

Energy

- Batteries have electricity inside them.
- Energy is a thing. This is a fuzzy notion, probably because of the way that we talk about newton-meters or joules. It is difficult to imagine an amount of an abstraction.
- Energy can be changed completely from one form to another (no energy losses).
- Things "use up" energy.
- Energy is confined to some particular origin, such as what we get from food or what the electric company sells.
- Energy is truly lost in many energy transformations.
- There is no relationship between matter and energy.
- If energy is conserved, why are we running out of it?

Forces and Fluids

- Objects float in water because they are lighter than water.
- Objects sink in water because they are heavier than water.
- Mass/volume/weight/heaviness/size/density may be perceived as equivalent.
- Wood floats and metal sinks.
- All objects containing air float.
- Liquids of high viscosity are also liquids with high density.
- Adhesion is the same as cohesion
- Heating air only makes it hotter.
- Pressure and force are synonymous.
- Pressure arises from moving fluids.
- Moving fluids contain higher pressure.
- Liquids rise in a straw because of "suction".
- Fluid pressure only acts downward.

Heat and Temperature

- Heat is a substance.
- Heat is not energy.
- Temperature is a property of a particular material or object. (Metal is

naturally cooler than plastic).

- The temperature of an object depends on its size.
- Heat and cold are different, rather than being opposite ends of a continuum.
- When temperature at boiling remains constant, something is "wrong".
- Boiling is the maximum temperature a substance can reach.
- Ice cannot change temperature.
- Objects of different temperature that are in contact with each other, or in contact with air at different temperature, do not necessarily move toward the same temperature.
- Heat only travels upward.
- Heat rises.
- The kinetic theory does not really explain heat transfer. (It is recited but not believed).
- Objects that readily become warm (conductors of heat) do not readily become cold.

Properties of Matter

- The bubbles in boiling water contain "air", "oxygen" or "nothing", rather than water vapor.
- Gases are not matter because most are invisible.
- Gases do not have mass.
- A "thick" liquid has a higher density than water.
- Mass and volume, which both describe an "amount of matter" are the same property.
- Air and oxygen are the same gas.
- Helium and hot air are the same gas.
- Expansion of matter is due to expansion of particles rather than to increased particle spacing.
- Particles of solids have no motion.
- Relative particle spacing among solids, liquids and gases (1:1:10) is incorrectly perceived and not generally related to the density of the states.
- Materials can only exhibit properties of one state of matter.
- Particles possess the same properties as the materials they compose. For example, atoms of copper are "orange and shiny", gas molecules are transparent, and solid molecules are hard.
- Melting/freezing and boiling/condensation are often understood only in terms of water.
- Particles are viewed as mini-versions of the substances they comprise.
- Particles are often misrepresented in sketches. No differentiation is made between atoms and molecules.
- Particles misrepresented and undifferentiated in concepts involving elements, compounds, mixtures, solutions and substances.
- Frequent disregard for particle conservation and orderliness when describing changes.
- Absence of conservation of particles during a chemical change.

- Chemical changes perceived as additive, rather than interactive. After chemical change the original substances are perceived as remaining, even though they are altered.
- Failure to perceive that individual substances and properties correspond to certain types of particles (i.e. formation of a new substance with new properties is seen as simple happening rather than as the result of particle rearrangement).

Measurement

- Measurement is only linear.
- Any quantity can be measured as accurately as you want.
- Children who have used measuring devices at home already know how to measure.
- The metric system is more accurate than the other measurement systems.
- The English system is easier to use than the metric system.
- You can only measure to the smallest unit shown on the measuring device.
- You should start at the end of the measuring device when measuring distance.
- Some objects cannot be measured because of their size or inaccessibility.
- The five senses are infallible.
- An object must be "touched" to measure it.
- Mass and weight are the same and they are equal at all times.
- Mass is a quantity that you get by weighing an object.
- Mass and volume are the same.
- Heat and temperature are the same.
- Heat is a substance.
- Cold is the opposite of heat and is a different substance.
- Surface area is a concept used only in mathematics classes.
- You cannot measure the volume of some objects because they do not have "regular" lengths, widths, or heights.
- An objects' volume is greater in water than in air.
- The density of an object depends only on its volume.
- Density for a given volume is always the same.
- The density of two samples of the same substance with different volumes or shapes cannot be the same.

Appendix C

Appendix D

http://www.ascd.org/readingroom/books/brooks99book.html#chap1

An excerpt from:

In Search of Understanding: The Case for Constructivist Classrooms
Revised Edition, 1999
by Jacqueline Grennon Brooks and Martin G. Brooks

The Construction of Understanding

It sounds like a simple proposition: we construct our own understandings of the world in which we live. We search for tools to help us understand our experiences. To do so is human nature. Our experiences lead us to conclude that some people are generous and other people are cheap of spirit, that representational government either works or doesn't, that fire burns us if we get too close, that rubber balls usually bounce, that most people enjoy compliments, and that cubes have six sides. These are some of the hundreds of thousands of understandings, some more complex than others, that we construct through reflection upon our interactions with objects and ideas.

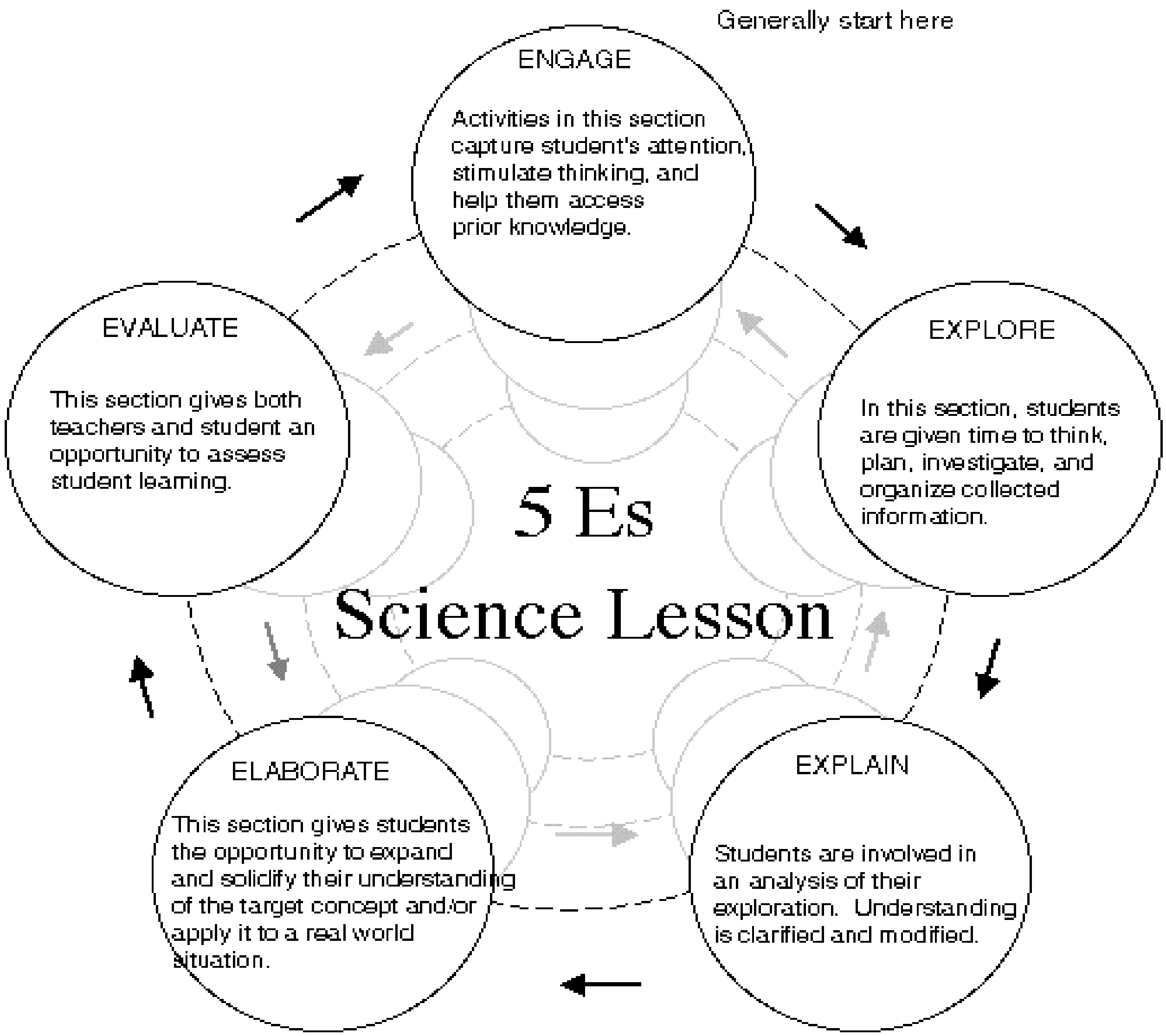
Each of us makes sense of our world by synthesizing new experiences into what we have previously come to understand. Often, we encounter an object, an idea, a relationship, or a phenomenon that doesn't quite make sense to us. When confronted with such initially discrepant data or perceptions, we either interpret what we see to conform to our present set of rules for explaining and ordering our world, or we generate a new set of rules that better accounts for what we perceive to be occurring. Either way, our perceptions and rules are constantly engaged in a grand dance that shapes our understandings.

Consider, for example, a young girl whose only experiences with water have been in a bathtub and a swimming pool. She experiences water as calm, moving only in response to the movements she makes. Now think of this same child's first encounter with an ocean beach. She experiences the waves swelling and crashing onto the shore, whitecaps appearing then suddenly vanishing, and the ocean itself rolling and pitching in a regular rhythm. When some of the water seeps into her mouth, the taste is entirely different from her prior experiences with the taste of water. She is confronted with a different experience of water, one that does not conform to her prior understanding. She must either actively construct a different understanding of water to accommodate her new experiences or ignore the new information and retain her original understanding. This, according to Piaget and Inhelder (1971), occurs because knowledge comes neither from the subject nor the object, but from the unity of the two. In this instance, the interactions of the child with the water, and the child's reflections on those interactions, will in all likelihood lead to structural changes in the way she thinks about water. Fosnot (in press) states it this way: "Learning is not discovering more, but interpreting through a different scheme or structure."

As human beings, we experience various aspects of the world, such as the beach, at different periods of development, and are thus able to construct more complex understandings. The young child in this example now knows that the taste of seawater is unpleasant. As she grows, she might understand that it tastes salty. As a teenager, she might understand the chemical concept of salinity. At some point in her development, she might examine how salt solutions conduct electricity or how the power of the tides can be harnessed as a source of usable energy. Each of these understandings will result from increased complexity in her thinking. Each new construction will depend upon her cognitive abilities to accommodate discrepant data and perceptions and her fund of experiences at the time.

Appendix D

A Learning Cycle for Constructivist Teaching…The 5 E's



Appendix D

THE FIVE E's

One instructional mode that supports a constructivist approach to learning is known as "5 E's". The five Es are engagement, exploration, explanation, elaboration, and evaluation. Below is a brief description of each of the phases of this instructional model. Usually you begin the study of a concept with the engagement phase. However, depending on the concepts that have been and will be investigated, you might begin at almost any point.

Engagement
This phase is design to grab a student's interest. An object, situation or problem that relates to the student's world is presented with an authentic question, problem description or an interactive scenario. This phase is designed to lead into the exploration tasks. The role of the teacher in this phase is to present the situation or problem and identify the task. A student's current understanding of concepts is elicited. If this phase is successful, students are motivated to continue to the exploration phase.

Exploration
Exploration activities are designed to provide students with concrete experiences upon which they continue to discover concepts, and learn new processes and skills. It brings answers to students, and if successful, satisfaction. During this exploration phase, students need time to explore objects, events, and/or situations. They gather data to help them develop relationships, construct mental pictures, observe patterns, and question preconceptions. The teacher facilitates the exploration and coaches students from the sidelines. The teacher answers student questions and helps them in restructuring their knowledge. At the end of this phase, students should be prepared to explain what they have discovered.

Explanation
This is the phase in which students should "see the light". The concepts, processes, and skills to which they have been exposed become clear. The learning begins to internalize. During the explanation phase, students and teachers should be able to reach the use of a common language to discuss the discoveries students have made. The teacher's role is to ask students to summarize what has happened in their own words. Then the teacher introduces scientific terms to describe the results and concepts. Explanation often gives order to the earlier phases and should lead quickly to the ability to elaborate on what has been learned.

Elaboration
This phase is designed to provide students with a chance to take what they have learned and apply or extend the concepts, processes, and skills. Often, elaboration activities are interdisciplinary and may involve writing, mathematics or social studies. When students can clearly connect the early explorations with explanations, and the concepts with the observations, internal learning has occurred. They are ready for evaluation of their work.

Evaluation
Students need to receive feedback on whether their explanations have been adequate. Informal evaluations occur all during the learning task, but a more formal evaluation should occur after the elaboration phase. Students should evaluate their own work and understanding, as well as be evaluated by the teacher. Authentic assessment techniques can be employed to give meaningful input on their individual work or any group work in which they participated.

Appendix D