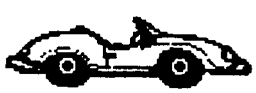
Mod Chem Bk
+ units for Honors!

What is Chemistry?

Or better yet, "What in the World Isn't Chemistry?"

Unit 1



Unit 1: What is Chemistry?

I. **Chemistry** is the study of the composition of substances and the changes that they undergo.

There are 5 major divisions of chemistry:

1) Organic chemistry - is the study of all substances that contain carbon
2) Inorganic chemistry - is the study of substances without carbon
3) Analytical chemistry - deals with determining the composition of substances
4) Physical chemistry - determines the theoretical basis for chemical behavior
5) Biochemistry - studies the composition and changes within living organisms

A chemist is trained to convert materials into substances that will be suitable substitutes for what is in short supply, such as scarce resources. She(He) has the experience to know how to change one substance into another with different properties. Technology is an integral part of the chemical industry that produces huge quantities of substances each year to meet the needs of today's society.

II. Altering the environment.

People have been altering their environment since the time of man to provide shelter and grow more food. Facts are needed to make intelligent choices about what long-term effects may occur from these alterations. The way facts are used involves value judgments. Science (learning facts) is neither "good" nor "bad". Benefits versus risks ("trade-offs") must be considered carefully before any alterations should be made.

Problem:
Discuss the use of a new chemical as a pesticide to prevent the damage caused by insects and disease to the food crops in our area. What facts do you need to know before you decide to use the chemical? At what point do the risks outweigh the benefits? Who should make the final decision about whether or not the pesticide will be used? Summarize your opinions below:

III. The Scientific Method

The Scientific Method is a formal approach to analyze and solve problems which incorporates:

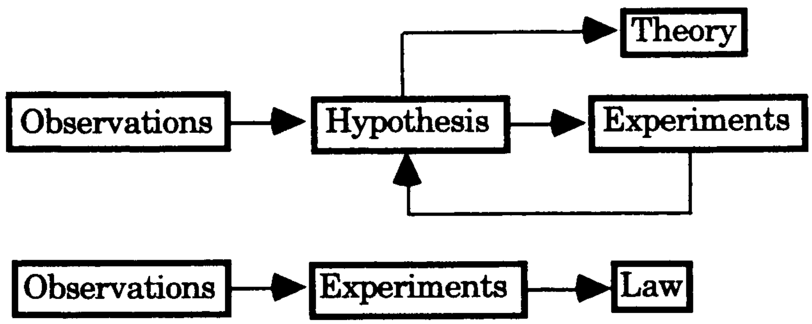
1) Observations: recorded facts about a natural phenomenon
2) Hypothesis: a descriptive model for the observations made
3) Experiments: one or more conditions are controlled to test the hypothesis

   Data: the observations which are recorded from the experiment
   If the hypothesis does not fit the data, it is refined and tested by further experimentation.

4) Theory: thoroughly tested hypothesis that explains why experiments give certain results
A theory can never be proved, but it provides generalizations that can be used to predict the behavior of natural systems under certain circumstances.

5) Law: a concise statement that summarizes the results of a wide variety of observations and experiments to describe a natural phenomenon, often by a simple mathematical relationship. For example, density = mass/volume

Outline of the Scientific Method:



A well designed experiment would contain the following components: 1)a clear and concise problem(hypothesis) to be solved 2)a very specific procedure that will provide measurable results when certain conditions are controlled 3)the results of the experimentation recorded as descriptions, measurements or calculated mathematical relationships and 4)your conclusions about the significance of the results....can you solve the initial problem? does your hypothesis fit the results? what modifications of the hypothesis or further experimentation is required?

Problems:
Classify as an example of an experiment, hypothesis, theory, or law.

a. The ashes from a campfire weigh less than the wood which was burned. Therefore mass is destroyed (or lost) when wood is burned.

b. An object at rest tends to remain at rest.

c. Water boils at 100°C on your kitchen stove and in the chemistry lab.

d. All matter is composed of atoms, which themselves are composed of protons, electrons, and neutrons.

IV. Alchemy and the Birth of Chemistry:

- Alchemists developed many experimental procedures and lab apparatus while attempting to change common metals into gold; primarily used trial and error methods.

- Roger Bacon suggested using observation and experimentation rather than pure logic to analyze and explain natural phenomena.

- Robert Boyle emphasized using experiments to test ideas that were obtained by reason.
- Antoine Lavoisier made precise measurements of mass changes during chemical reactions, and he is called the founder of modern chemistry because of his emphasis on careful experimental measurements.

V. Properties of Matter

Matter is defined as anything which takes up space(has a volume) and has mass.
examples: ice, water, steam, air, sugar, vegetable oil

1) **Mass** is the amount of matter that an object contains. It has a constant value that is measured by using a balance(like a see-saw) to compare the mass of the object to the known mass of standard objects.
2) **Weight** is the amount of gravitational attraction for an object, and weight will vary depending on location.
   For example, an object that has both a mass and weight of 12 kilograms on Earth will have a mass of 12 kilograms on the Moon, but a weight of only 2 kilograms due to the weaker gravity of the Moon(about 1/6th the gravity of the Earth).
3) All matter has a property called **inertia**, which is the amount of resistance the object has to any change in its motion or position. Inertia increases with the objects mass and velocity.

A **pure substance**, such as table salt, is a particular kind of matter that has a uniform and definite composition. All samples of a pure substance have identical properties.

Physical properties of a substance are characteristics that can be observed or measured without changing the substance's composition.
examples: melting point, boiling point, density, color, odor, hardness, mass

Problem:

List the physical properties of water that could be used to identify it in the laboratory.

VI. States of Matter

Depending on the experimental conditions of temperature and atmospheric pressure, matter can exist in three different physical states as either a solid, liquid or gas. Each state has certain characteristics.

Solids have a fixed mass and volume with a high density which makes them noncompressible. The particles are held together in a rigid structure with a definite arrangement. As the temperature of a solid increases, the particles expand very slightly until melting occurs to form a liquid.

Liquids also have a fixed mass and volume and are noncompressible, but the particles can flow and tumble over each other. They are generally less dense than solids, and they take the shape of their container beneath the surface of the liquid. As the temperature increases, molecules evaporate by breaking away from the surface of the liquid to form a gas. At the boiling point, all of the molecules start to separate from each other.

Gases expand without limit to fill any space and are easily compressible. They have a definite mass, but no definite shape or volume. Gases take the shape and volume of their containers with the particles randomly bouncing off each other and the walls of the container. The term vapor is used to describe a substance that generally exists as a solid or liquid at room temperature, but is presently in the gaseous state.

Problem:
What is the physical state of each of the following at room temperature?
a. silver b. gasoline c. helium d. paraffin wax e. rubbing alcohol

VII. Physical Changes

A **physical change** will alter the substance without changing its composition. It may involve only a change in appearance, such as tearing a piece of paper or bending a nail, or it may involve a change in state, such as melting an ice cube or dissolving sugar in a cup of tea.

Problem:
List ten verbs that indicate a physical change in matter occurs.

VIII. Mixtures

Mixtures consist of a physical blend of two or more substances. They may have a variable composition, but each component retains its own unique characteristic properties. The differences in melting and boiling points, solubilities and density can be used to separate the mixtures in the laboratory.

Heterogeneous mixtures are not uniform in composition, and distinct phases or layers are usually visible to the naked eye.
Examples: soil, sand in water, an ice cream soda, oil and water

Heterogeneous mixtures can be separated in the lab by pouring off less dense layers from the top(decantation) or using porous filters to remove larger particles that settle out.

Homogeneous mixtures (called **solutions**) has a completely uniform composition because its components are evenly distributed throughout the sample. It appears to have only one phase of matter.
Examples: ocean salt water, the air, brass metal

A solution consists of a smaller amount of one substance called the **solute** which is dispersed in the larger amount of another substance called the **solvent**. For the ocean salt water example, the solute is the solid salt that has been dissolved in the larger volume of liquid water as the solvent. Solutions can be separated in the lab by evaporation to recover dissolved solids or by distillation which boils off a liquid to produce its vapor which is recondensed into the liquid.

An example of each type of solution consisting of the various solute/solvent pairs is given in the chart below:

Solute/Solvent	**example**
gas-gas	oxygen gas dispersed in the atmosphere
liquid-gas	water vapor in air(steam)
gas-liquid	carbon dioxide in water (carbonated soda water)
liquid-liquid	ethanol(ethyl alcohol) in water
solid-liquid	table sugar in water
solid-solid	zinc in copper(brass)

Problem:
Describe how you would separate a mixture of iron filings, sawdust, sand and salt to recover each of the 4 components.

IX. Compounds and Elements

When mixtures are separated in the laboratory, pure substances are obtained. These substances can be divided into two groups, compounds or elements.

Compounds are substances that can be broken down into simpler substances(elements) only by chemical reactions. During a **chemical reaction** new substances with different properties are formed. The component elements of a compound are always present in the same proportions, and every compound has its own unique set of properties.

Elements are the simplest forms of matter that can exist under normal laboratory conditions. They cannot be broken down further by chemical reactions and are the building blocks of all other substances. The smallest unit of an element that still retains its characteristic properties is called an **atom**. As of this writing, there are 110 known elements, of which 91 occur naturally.

Chemists further classify elements and compounds into groups based upon similar properties. For example, elements can be described as being either ***metallic*** or ***nonmetallic***. Most metals are shiny, flexible solids that are good conductors of heat and electricity. Nonmetals exist as gases, or easily vaporized liquids and solids, that have a dull appearance and absorb electricity. Many compounds can be grouped as being either an ***acid*** or a ***base***. Acids increase the concentration of hydrogen ions(H^{1+}) in solution, dissolve metals, taste sour and have low pH values. Bases increase the concentration of hydroxide ions(OH^{1-}) in solution, react with acids to form water, taste bitter and have high pH values. These classifications help chemists to predict the properties and behaviors of the millions of substances that exist.

Problem:
When a solution of table sugar dissolved in water is heated to complete dryness, the following changes will occur. First, all the water will boil off leaving solid sugar behind in the container. Then the sugar will begin to decompose into black carbon ash and more water vapor will be formed. Identify all of the materials as either elements, compounds or mixtures and state whether each change is a physical change or a chemical reaction.

X. Protons, Electrons and Neutrons

Scientists have discovered that all atoms are made up of three fundamental subatomic particles called protons, electrons and neutrons. The number of protons in an atom is known as the **atomic number**(represented by the symbol Z), and it determines the identity of the element. For example, all atoms of hydrogen contain 1 proton. An atom with 2 protons would be classified as a helium atom. A **proton** has a positive 1 electrical charge and is found in the center of the atom in an area called the **nucleus**.

Electrons are very lightweight particles(about 1/1800 the mass of a proton) that have a negative 1 electrical charge. They spin around the nucleus of the atom, creating a cloud region of negative charge. This forms a force field that gives the atom its volume and prevents other atoms from entering its space. Since atoms are electrically neutral, the number of protons equals the number of electrons.

A **neutron** is another particle in the nucleus that has about the same mass as a proton but no electrical charge. Most of the mass of an atom is due to the total number of protons and neutrons. This sum is called the **mass number**(symbolized as A). The number of neutrons in the atoms of a particular element can vary, giving them slightly different masses. These different atoms of the same element are called **isotopes**. The atomic mass(sometimes called atomic weight) value on the periodic table is actually a weighted average of the masses of all of the atoms of an element.

Sample Problem:
Lithum has an atomic number of 3 and an atomic mass of 6.941. How many protons, electrons and neutrons are in an average atom of lithium?
Answer: There are 3 protons, 3 electrons and 4 neutrons in the average lithium atom.

Problems:

1. What is the name of the element that has 26 protons?

2. What is the name of the element whose atom has 53 electrons?

3. How many neutrons are in the average atom of uranium?

4. How many protons, electrons and neutrons are in an isotope of the element with an atomic number of 79 and a mass number of 200?

XI. Chemical Symbols

Each element is represented by a **chemical symbol**. For most elements the symbol is the first letter or two of the element. The first letter of the symbol is always capitalized and a second letter, if needed, must be lower case.
Examples: hydrogen H helium He lithium Li carbon C

When several elements begin with the same letter, a second, pre-dominant letter from the name is selected to identify the different elements.
Examples: carbon C calcium Ca chlorine Cl californium Cf

Some element symbols were derived from older Latin names, instead of their modern names, so their symbols are not consistent with their common names.
Examples: sodium Na (natrium) potassium K (kalium) gold Au (aurum) lead Pb (plumbum) iron Fe (ferrum) tungsten W (wolfram; not Latin)

Chemical symbols are used to write **chemical formulas** that represent the proportions of various elements in a compound. Subscript numbers indicate the ratios of each element in the compound.

Examples: water is made of 2 hydrogen atoms and 1 oxygen atom; the formula is H_2O and sugar, $C_{12}H_{22}O_{11}$, has 12 carbon atoms, 22 hydrogen atoms and 11 oxygen atoms per molecule.

If some table salt, NaCl, is blended together with the sugar, the mixture would be represented by both formulas joined with an addition sign; $NaCl + C_{12}H_{22}O_{11}$,

Problem:

Classify each of the following as an element, compound or mixture.

a. argon b. ethyl alcohol (C_2H_5OH) c. grape juice d. cheese e. zinc

In order to be able to speak the language of chemistry, it is necessary to know the names and symbols of the most common elements.

DIRECTIONS: MEMORIZE THE FOLLOWING SYMBOLS AND NAMES.

1. aluminum	Al	20. gold	Au	39. plutonium	Pu
2. americium	Am	21. helium	He	40. potassium	K
3. argon	Ar	22. hydrogen	H	41. radium	Ra
4. arsenic	As	23. iodine	I	42. radon	Rn
5. barium	Ba	24. iron	Fe	43. rubidium	Rb
6. beryllium	Be	25. krypton	Kr	44. silicon	Si
7. bismuth	Bi	26. lead	Pb	45. silver	Ag
8. boron	B	27. lithium	Li	46. sodium	Na
9. bromine	Br	28. magnesium	Mg	47. strontium	Sr
10. cadmium	Cd	29. manganese	Mn	48. sulfur	S
11. calcium	Ca	30. mercury	Hg	49. thallium	Tl
12. carbon	C	31. molybdenum	Mo	50. tin	Sn
13. cesium	Cs	32. neon	Ne	51. titanium	Ti
14. chlorine	Cl	33. nickel	Ni	52. tungsten	W
15. chromium	Cr	34. nitrogen	N	53. uranium	U
16. cobalt	Co	35. oxygen	O	54. vanadium	V
17. copper	Cu	36. palladium	Pd	55. xenon	Xe
18. fluorine	F	37. phosphorus	P	56. zinc	Zn
19. francium	Fr	38. platinum	Pt		

XII. Chemical Reactions

In a chemical reaction, one or more of the starting substances, called **reactants**, are changed into new substances called **products**. Chemists use a shorthand expression to show the change.

During a chemical reaction, several occurrences indicate that a chemical change is taking place. These include:

1) production of energy in the form of heat, light or electricity
2) change in color or odor
3) production of a gas or a solid(a precipitate) from a solution
4) the change cannot be easily reversed.

Chemical properties describe the ability to and behavior of a substance when it is undergoing a chemical reaction.
Examples: rusting is a chemical property of iron; burning is a chemical property of wood

Problem:
Classify the following changes as physical or chemical changes.

a. food spoils   b. water boils   c. a nail rusts   d. a firefly emits light

e. bread is baked   f. oil is pumped out of a well   g. sugar dissolves in water

h. a snowflake melts   i. paper burns   j. a battery starts a car engine

XIII. Conservation of mass and energy

During any physical or chemical reaction, the quantity of matter remains unchanged. Therefore the total mass of the reactants must equal the total mass of the products. This is the Law of Conservation of Mass.

All physical and chemical changes in matter also involve changes in the energy of the system. Like mass, energy is also conserved and neither created nor destroyed. It may only be converted from one form to another. Einstein related mass and energy by the equation, $E=mc^2$, where c is the speed of light.

Energy is defined as the capacity for doing work.

From physics, **work** equals a force applied through a distance.

Energy may exist in several forms including: chemical, nuclear, radiant, electrical, mechanical and thermal energies.

Potential energy is stored energy, or the energy of position, that provides the ability to do work. *Example:* A piece of coal has potental energy that can be released when it burns.

Kinetic energy is the energy of motion that occurs when work is actually being done. *Example:* Expanding gases produced by the explosions inside your car's engine drive the pistons to make the wheels turn.

Mathematically, $KE = 1/2\ mv^2$

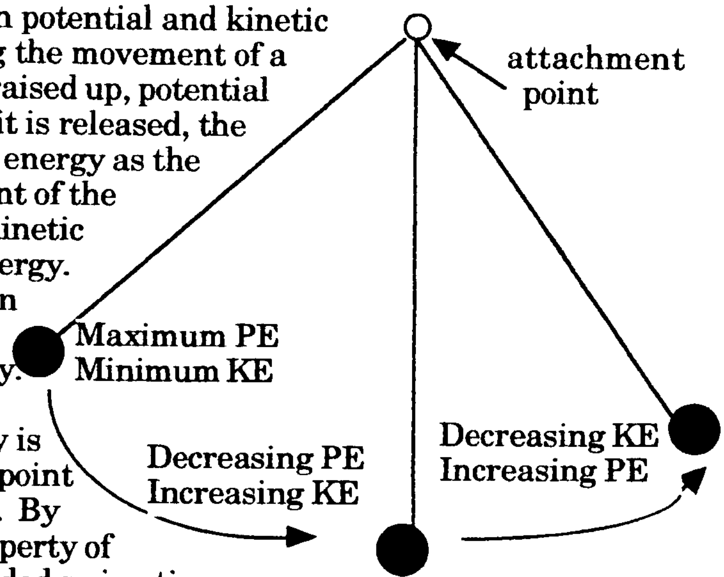
kinetic energy equals 1/2 the mass of an object multiplied by its velocity squared.

Temperature provides a measure of the average kinetic energy of the particles in a material. The higher the temperature, the faster the velocity of the particles.
Example: Boiling water at 100°C has more kinetic energy than water molecules at a room temperature of 20°C. Drops of food coloring will mix faster in the hot water.

Heat is a measure of the amount of energy that can be transferred from a warm object to a cooler object because of a difference in temperature. The amount of heat an object has depends on *both* the temperature of the object and its mass. Heat content increases with increasing temperature and/or greater mass.

Example: One drop of boiling water at 100°C has less heat energy than 1 liter of water molecules at 20°C.

A summary of the transitions between potential and kinetic energy can be understood by studying the movement of a large pendulum. When the weight is raised up, potential energy is added to the system. Once it is released, the potential energy is changed to kinetic energy as the weight gains speed. At the lowest point of the swing, the weight has the maximum kinetic energy and the minimum potential energy. On the upswing, the weight slows down as it gains height, and kinetic energy is converted back into potential energy. The height of each successive swing becomes less and less, as some energy is lost due to friction at the attachment point and from collisions with air molecules. By substituting a heavier weight, the property of inertia can also be shown by the extended swing time.



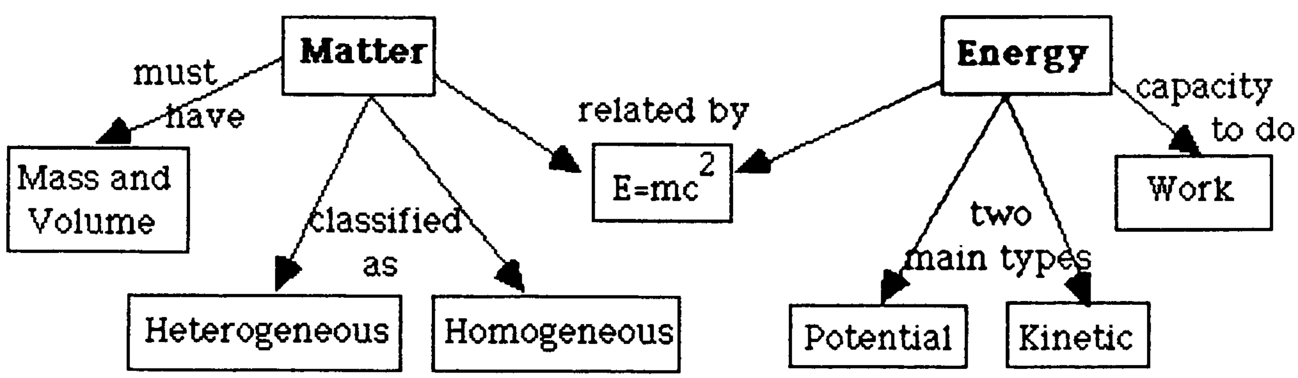
Problems:

1. Describe all of the interconversions of energy that occurs when a candle burns. How many different forms of energy are involved?

2. Describe the changes in temperature and heat energy that occur when ice cubes are added to a glass containing warm soda.

Review Problem:

Prepare a concept map to summarize the key concepts of this unit, starting with Matter and Energy as the primary headings. Include key terms and concepts as subheadings and connect them with linking phrases. Follow the pattern started in the map below:



Unit 1 Objectives

Having studied the unit notes and done the problems, you should be able to:

1. Define Chemistry and state the role of chemists.

2. Show how the terms experiment, hypothesis, theory, and law fit into the scientific method.

3. Describe the characteristic properties of all matter..

4. Using a microscopic model, describe how solids, liquids and gases differ and how a substance changes from one physical state to another.

5. Indicate whether an observed change in matter is chemical or physical and state reasons for your conclusions.

6. Describe at least 3 ways that a mixture can be separated, and name the property of the substance in the mixture that allows it to be separated.

7. Describe the difference between the proportions of elements in a compound and in a mixture, and the microscopic difference between a heterogeneous mixture, a solution and a pure substance.

8. State how to determine whether a pure substance is an element or a compound.

9. Determine the number of protons, neutrons and electrons in an atom based on the element's name or the atomic number and mass number(or atomic mass).

10. Write the symbols of the common elements, and write the names of common elements given their symbols.

11. Define a chemical reaction and describe the evidence that indicates a chemical change has occurred..

12. Distinguish between potential energy and kinetic energy.

13. Compare the differences in temperature and heat energy of two objects.

14. Give an example of the conversion of one form of energy to another.

15. State the law of conservation of mass and/or energy.

16. Students will be aware of and practice the American Chemical Society safety procedures for handling chemicals.