Lecture Notes: Chem 101 Chaps. 1+2 Week 1 (pgs. 1 - 42 of Text)

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What is chemistry?

1. Explains: Rusting faster with salt, fire 2. Predicts: Calcium will react with water; copper will not. **3. Helps**: Medicines, chemotherapy, vitamins and minerals, plastics, determining risk/health factors involved in certain habits like drinking or smoking, being exposed to radon **4. Hurts**: explosives, Alfred Nobel, *Poisons*: arsenic, sarin, cyanide **5. Fun**: sugar fuel rockets, fireworks, glow-in-dark slime/gak, vinegar and baking soda volcanoes

Some definitions of chemistry (p. 2)**:** Chemistry has been called the "central science." This is due to it being used as a focus point for all other sciences. **Example**: Biology using chemistry to explain the complex chemical reactions needed for life. ***Chemistry is the study of matter and its changes. It is concerned with the building blocks of all we see and even cannot see (most gases for instance) around us and how these interact with each other.***

Matter: Anything that has mass (a measure of the amount of matter in an object) and takes up space.

The Scientific Method (p. 4): Can be considered as a process people (science related work in particular) use to solve problems.

5 main steps: 1. Clearly define the problem, 2. Gather and collect information, 3. Form a hypothesis (educated guess), 4. Test the hypothesis, 5. Draw a conclusion based upon the results

NOTE: The steps may not always be followed in this order and after a conclusion is drawn, new experiments may be needed to explain other events that were discovered. EXAMPLE: The discovery that mold prohibited growth of certain bacteria near it was a mistake resulting from the attempt to solve another problem (studying growth of bacteria).

Theory vs. Law: A theory is a probable (possible) explanation for a series of events or circumstances. EXAMPLE: *The crystal-field theory* attempting to

explain the different colors transition metal ions have in aqueous solutions.

A Law is a theory that has been proven over and over again to the point that no exceptions under the given conditions are found. A law is a fundamental understanding and can provide the foundation for new studies.

EXAMPLE: The *second law of thermodynamics* stating whenever an event occurs spontaneously (on its own) in our universe, it is generally accompanied with an overall increase in entropy. Example: Objects fall in disarray or disorder not orderly such as what happens when dropping a stack of papers.

Physical States of Matter and Their Characteristics

Atoms vs. Molecules

Atoms are the smallest components of elements. Atoms of hydrogen, helium, lithium, beryllium, boron, etc. are each unique and have differing properties. Molecules are the smallest components of compounds. Molecules of water, carbon dioxide, chlorine (an element that exists as a molecule), sulfuric acid, etc. are the chemical combination of atoms into new substances.

See Table 1.2 pg. 8 for a summary of these physical states of matter.

a. Solids: ***definite shape and volume***; some are crystalline: They have repeating, 3-D geometric patterns. *Examples*: salt, quartz, sand, galvanized zinc

Not all are crystalline, some and be amorphous: have no repeating, definite shapes and form (random particle arrangement). *Examples*: glass, red phosphorus (on match box strips), sawdust

b. Liquids: ***Definite volume and indefinite shape***; Liquids take the shape of the container they are in and are not easily compressed (have their volume changed); they can also be poured.

c. Gases: ***indefinite volume and shape***; Gases like liquids, take the shape of their container; however, they are easily compressed as the particles are very spread out. (A gas is mostly empty space.) Gases can also diffuse or move from an area of greater concentration to an area of lesser concentration and can spread out through a container, room, building, even the atmosphere itself. *Example*: skunk spray, chemical name = but-2-en-1-thiol

See Table 1.1 pg. 8 for common substances that exist in one of the phases above.

d. Plasma: similar to a gas except it exists at only extreme temperatures that occur only in stars; plasma is a fully ionized (the atoms have lost all electrons) gas that has about the same amount of positive and negative ions. Plasma is the most abundant physical state of matter in the universe due to stars being so abundant.

Substances vs. Mixtures

Substances: Matter composed of the same material throughout. These materials can be **elements or compounds** and can be separated only by chemical means (such as burning or heating to decomposition or reaction with another substance).

Mixtures: Matter that is not the same throughout: can be separated by physical means such as filtering or using a magnet.

Mixtures can be *homogeneous* or *heterogeneous*. Another name for a homogeneous mixture is a *solution*. A solution is the same throughout (uniform), has the same phase (uniform distribution of materials as a solid (metal alloys such as brass), liquid (antifreeze), or gas (air)).

Heterogeneous: different substances that can be separated by physical means. In this type of mixture, the material substances composing it are easily distinguished. Examples: salads, pizza, orange juice with pulp

Chapter 2: Standards of Measurement pp. 15-40

Physical States of Matter and Their Characteristics

Mass vs. Weight: Weight depends on gravity; gravity depends on the mass of the object and its distance away from another object attracted to it. Example: We would weigh a lot less on Pluto because Pluto is less massive than Earth. If we could stand on the surface of the sun, we would weight a lot more than on Earth. **Mass of an object is constant and is not affected by gravity**.

Scientific Notation of Numbers

Important as it allows really big and really small measurements to be shortened.
Powers of 10 are used in place of some digits (usually zeroes).
RULES: If the number is less than (<) 1, a negative exponent is used with the 10; if the number is greater than (> 1), a positive exponent is used. OR If the decimal is moved to the right to give a number between 1 and 10 (but not including 10), the exponent used is negative and equals the number of spaces moved to acquire that number between 1 and 10; vice versa if positive exponent used.
Examples: **1000000**: This would be written as 1.0×10^{6} and **0.000001** as 1.0×10^{-6}

Taking Measurements Using Significant Figures

Measurements are numbers usually acquired by using certain equipment like a scale for weight or ruler for length or thermometer for temperature.

A **significant figure** is a number (digit) that contains all digits that are known to be certain PLUS one that is uncertain (not known for sure).

<u>Example</u>: A measurement on a scale of the mass of a paper clip is found to be 1.28 grams. The scale reads to the nearest hundredth place (the 0.08). The uncertain digit would be the hundredths place digit, so the number of significant figures would be 3 (the 1, the 2, and the 8). The thousandths place number would not be significant.

How do you know which digits are significant when looking at a number?

GUIDELINES FOR DETERMINING SIGNIFICANT FIGURES

The only number that is not always significant is 0! The numbers 1-9 are always significant!

Hence, a 0 is significant when,

1. It is between 2 non-zero digits. Ex.: **1.202** has 4 significant figures; **88909.12** has 7 sig figs.
2. It is at the end of a number that includes a decimal point. Here it is showing certainty and is significant. Ex.: **0.900** has 3 sig figs; **9.000** has 4; **1000.** has 4, **0.0090** has 2 because the first 3 zeros serve as place holders only!; **1000** has only 1 sig fig as the 3 zeros are place holders too.

2 Rounding Rules

1. If the last number in a measurement is 4 or less, round down so that last digit and all others to its right are dropped. Ex. **1.973** would be rounded to **1.97** with the 3 being dropped as it is <4.

2. If the number is 5 or higher, round up so that last digit and all others to the right are dropped and that last digit is increased by 1. Ex.: **1.975** rounds to **1.98** with 5 dropped.

Multiplication/Division Using Sig. Figs.

When multiplying or dividing measurements, the answer should be rounded to contain the same number of sig figs as the measurement that had the LEAST number of sig figs originally.

Ex. 10 x 12 =120, the answer, using sig figs, would be 100 as 10 only has one sig fig, so 120 was rounded to 100 as it is closer to it than 200. Ex. 12 ÷ 10 = 1.2, using sig figs answer is 1.

Addition/Subtraction Using Sig. Figs.

When adding or subtracting sig figs, answer should be rounded to contain the same number of sig figs as that measurement which has the least number of decimal places. Ex. 12.1 - 10.05 = 2.05 answer using sig figs would be 2.1 as 12.1 has fewest decimal places.

The Metric System or International System (SI = Systeme International, French)

This measurement scale is based on powers of 10 and uses prefixes to indicate which power. See Table 2.1 in text for prefixes.

Standard Units of Measure: Length = meter; Mass = kilogram; Temperature=Kelvin; Time =Second; Quantity of a substance = mole; Electric Current = ampere; Power = watts; etc.

SEE TABLE 2.2 ON P. 23.

Measurement of Length: Different in the past: SCIENTISTS USE METRIC PREFIXES WITH THE METER (m) AS THE BASE UNIT. Prefixes indicate the power of 10 for the number. Ex.: 0.1 m = 1 decimeter (deci- = tenth).
10 m = 1 dekameter (deka = 10)

FACTOR LABEL (Dimensional analysis) METHOD OF PROBLEM SOLVING

A method of problem solving in which units are converted from one unit to another. The goal is to get from given units that you know to unknown units that you want to find.

BASIC STEPS LISTED ARE ON P. 25.

For instance, you have a meter stick or ruler that reads in the English system only and you need units in a formula that have to be metric. So, you would need to convert from English units to metric units. Ex.: 12 in = ? m SEE APPENDIX III ON PG. A-16 OR INSIDE OF BACK COVER FOR CONVERSION FACTORS.
12 in. x 2.54 cm/1 in. x 1 m/ 100 cm: Units on top cancel with units on bottom to give desired unit: meters. **Some problems require only 1 fraction (conversion factor) others can require many. Look for the simplest route knowing the conversion factors you are given or already know.**

MEASUREMENTS OF MASS: Mass can have different units; we will be using the metric system unit of mass in most calculations: the kilogram (1000 g).
1 kg = 1000 g = 2.205 lb. = 10^6 mg. Using the dimensional analysis method, we can convert between one unit of mass to another.

MEASURMENTS OF VOLUME: Volume is the amount of space matter occupies. The metric unit of volume is the liter (L) and = 1 dm^3 = 1000 mL = 1000 cm^3 = 1.057 qt.

Volume can be measured using: graduated cylinders, volumetric flasks, burets, pipets, and syringes accurately.

MEASUREMENT OF TEMPERATURE: Temperature is a measure of the average kinetic energy of a substance or mixture.

a. **Fahrenheit scale: This is the temperature scale that we most commonly use. Scientists and other professionals usually do NOT use this scale; rather they use the Celsius or Kelvin scale (Astronomers may even use the Rankin scale.)**
    i. Devised by the German physicist Daniel Gabriel Fahrenheit around 1724.
    ii. He was the first scientist to use mercury rather than alcohol in a thermometer.
    iii. He assigned the zero (0) temperature value of his scale based upon the coldest temperature reading he could acquire using a mixture of salt, ice, and water.
    iv. Water freezes at 32^{o}F and boils at 212 ^{o}F (at sea level pressure).

b. **Celsius scale: This is one of the temperature scales often used by scientists.**
    i. **Devised by Swedish astronomer Anders Celsius in 1742.**
    ii. **The value of 0 temperature is assigned to the freezing point of water and 100 to the boiling point.**

c. **Kelvin or Absolute scale:** This is the temperature most commonly used by scientists.

   i. Named after Sir William Thomson or Lord Kelvin

The value of 0 is assigned as the coldest possible temperature that can be reached and the freezing point of water is 273.15 K on his scale. Note: The degree symbol is not used.

To convert from one temperature to another, use the following equations:

1. Celsius to Fahrenheit: $^{o}F = (9/5* {}^{o}C) +32$ (Do what is in parentheses first!)

2. Fahrenheit to Celsius: $^{o}C = 5/9(^{o}F-32)$

3. Celsius to Kelvin: $K = {}^{o}C + 273.15$

4. Kelvin to Celsius: $^{o}C = K - 273.15$

MEASUREMENT OF DENSITY: Density is the mass per unit volume of a substance. Its units are mass/volume. **Exs**.: g/ml; kg/L Density, like volume, depends on temperature. It is a piece of data that can be used to identify a substance (See Table 2.5 p. 38).

Specific Gravity: The relation of a substance's density to that of water; a ratio of density of substance to density of water. Sp. Gr. = density of substance/1.0 g/mL. Units cancel and since dividing by 1, Sp. Gr. = density of substance without any units.