Lecture Notes: Chem 110 Chapter 1

Overview of syllabus: **e-mail**: vtc101@psu.edu

What is chemistry?

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 most other sciences.

Example: Paleontology using chemistry to determine if DNA is still present in fossils.

Common Definition: Chemistry is the study of matter and its changes. It is concerned with the building blocks of all we see and even cannot see around us and how these interact with each other to make up matter.

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

The Scientific Method (pg 15): Can be considered as a process used by scientists and others to aid in solving 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 or the hypothesis may need scrapped.

EXAMPLE: The discovery that mold prohibited growth of 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 VSEPR theory* attempting to explain the shapes molecules possess when chemical bonding occurs.

A **Law** is the result of an event being verified repeatedly enough that it is taken as being valid under the given circumstances. 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 (Disorder).
Example of 2nd Law: When dropped, objects tend to fall in disarray or disorder as is the case with a tray of dishes or stack of papers.

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. Mass of an object is constant and is NOT affected by gravity.

Properties of Matter

2 Main Types of Properties

a. **Physical properties:** Characteristics of a substance (element or compound) that do not change and are useful in identifying the substance.

Examples: weight, mass, height, volume, color, density, specific gravity, size, shape, temperature, melting (freezing) point, boiling point, odor, etc. are all PHYSICAL properties

b. **Chemical properties:** Characteristics of a substance that indicate how the substance interacts or changes in the presence of other substances.

Examples: flammable, explosive, bleaching, oxidizing, corrosive, caustic, unstable, reacts slowly, reacts vigorously, color changing, or polymerizing are all terms that relate the possible CHEMICAL properties of a substance.

NOTE: Properties can be either Extensive (depend upon the amount) such as the weight of an object OR

Intensive (do NOT depend upon the amount present) such as the density of a substance

PHYSICAL VS. CHEMICAL CHANGE

Physical changes are those that do NOT result in a change in the identity of a substance; these changes simply result in a change in physical state (solid, liquid, gas, or plasma).

Examples: ice melting, water boiling, water condensing, ice freezing, alcohol evaporating, iodine subliming (going directly from solid to gas)

Chemical changes are those that DO result in a change in the identity of the substance; these changes result in a new substance forming.

Examples: paper burning to form carbon dioxide and water vapor (combustion); a marshmallow charring to produce carbon and water vapor (decomposition); silver tarnishing in air to produce silver (I) sulfide (synthesis); chlorine replacing iodine from sea water (single replacement); hard water stains forming from calcium and magnesium salts depositing on a bath tub (double displacement)

Taking Measurements with Application of 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).

Example: A measurement on a scale (balance) of the mass of a paper clip is found to be 1.284 grams. The scale reads accurately to the nearest hundredth place (the 0.08). The uncertain digit would be the thousandths place digit (the 0.004), so the number of significant figures would be four (the 1, the 2, the 8, and the 4).

***KEY QUESTION*: 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 through 9 are always significant!

Hence, we are only concerned with the zero(s) present in a number; the following rules apply: The zero is…

1. between 2 non-zero digits. Exs.: **1.202** has four significant figures; **88909.12** has seven sig. figs.
2. at the end of a number that includes a decimal point. Here it is showing certainty and is significant.

   Exs.: **0.900** has three sig. figs.; **9.000** has four; **1000.** has four, **0.0090** has two because the first 3 zeros serve as place holders only!; **1000** has only one sig. fig. as the 3 zeros are place holders too. *NOTE: In 1000. the decimal point makes 0's significant.*

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 since $5 \geq 5$.

Scientific Notation of Numbers

An important method of representing really big and really small measurements as it allows them to be shortened and easier to write.

Powers of 10 are used in place of some digits (usually zeroes).

$$\text{FORM: Number} \times 10^{\text{some power (+ or -)}}$$

RULES: If the number is less than 1, a negative exponent is used with the power of 10; if the number is >10, a positive exponent (1 or higher) is used; if the number is between one and ten, 0 is used as the exponent.

Examples: **1,500,000**: This would be written as 1.5×10^6 and **0.0000015** as 1.5×10^{-6}.

1.5 would be written as 1.5×10^0 since 1.5 is between 1 and 10.

NOTE: 1,500.0 would be written as 1.5000×10^3. Zeroes are shown due to accuracy being represented in the 1,500.0; hence, these zeroes are significant.

Multiplication/Division Rounding Rules 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 has the LEAST number of sig. figs. originally.

Ex. 1 **170 ÷ 100 = 1.7**, the answer, using sig. figs., would be **2** as 100 only has one sig. fig. while 170 has two, so 1.7 was rounded to 2 (least number of sig. figs. involved is one).

Ex. 2 15.030 x 0.02 x 903 = 271.4418

↑ ↑ ↑

5 1 3

Using sig. fig. rounding rules, the answer is 300.

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. 12.108 - 10.05 = 2.058; answer using sig. figs. would be 2.06 as 10.05 has fewer decimal places than 12.108.

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 1.5 in text for prefixes. ***Be aware of these prefixes and their meaning!***

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

SEE TABLE 1. 4 ON PG. 15.

Measurement of Length: Different in the past: Now, USE METRIC PREFIXES WITH THE METER (m NOT m- which represents milli-) AS THE BASE UNIT.

Prefixes in metric system indicate the power of 10 for the number. **Ex.**: 0.01 m = 1 centimeter (centi- = hundredth); 100 m = 1 hectometer (hecto- = 100)

FACTOR LABEL (Dimensional analysis) METHOD OF PROBLEM SOLVING

This is a method of problem solving in which units are converted from one (given) to another (desired). The goal is to get from given units that you know to unknown units that you want to find or desire.

See PGS. 25 - 29 in text for more information.

For instance, you have a yard stick or ruler that reads in the English system only and you need metric units to use in a formula for calculating volume. So, you would need to convert from English units to metric units.

SEE INSIDE BACK COVER OF TEXT FOR CONVERSION FACTORS.

Ex.: 16.5 in = ? m

$$16.5\ \cancel{\text{in}} \times \frac{2.54\ \cancel{cm}}{1\ \cancel{in}} \times \frac{1\ m}{100\ \cancel{cm}} = 0.4191\ \text{m} \rightarrow 0.419\ \text{m}$$

Answer is rounded to three sig. figs. due to given value (16.5) having three sig.

figs.

<u>DO NOT USE CONVERSION FACTORS FOR DETERMINING ROUNDING OF SIG. FIGS.!</u>

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. <u>NOTE: You will be given conversion factors for test.</u>**

<u>MEASUREMENTS OF MASS</u>: Mass can have different units; we will be using the metric system unit of mass in most calculations: **the gram** (1 g).
1 kg = 1000 g = 2.205 lb = 0.4536 g = 10^6 mg. Using the dimensional analysis method, we can convert between one unit of mass to another.

<u>MEASURMENTS OF VOLUME</u>: Volume is a measure of 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 = 0.2643 gal
Volume can be measured accurately using: graduated cylinders, volumetric flasks, burets, pipets, and syringes **NOT** beakers or flasks.

<u>MEASUREMENTS OF TEMPERATURE</u>: Temperature is a measure of the average kinetic energy of a substance or mixture.

a. **<u>Fahrenheit scale</u>: This is the temperature scale that we most commonly use in U.S. 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 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 is assigned to the freezing point of water and 100 to the boiling point of water.**

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.

<u>Note</u>: The degree symbol is not used with this scale.

<u>To convert from one temperature to another, use the following equations:</u>

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

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

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

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

5. Kelvin to Fahrenheit: $^{o}F = \frac{9}{5}(K - 273.15) + 32$

6. Fahrenheit to Kelvin: $K = \frac{5}{9}(^{o}F - 32) + 273.15$

MEASUREMENTS 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 physical property that can be used to identify a substance .

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 without any units.