Instructor: Rob O’Brien (Sci 163) 807-9569 email: Rob.OBrien@ubc.ca

Course Description: Fundamentals of structural chemistry; descriptive chemistry of main-group elements, with industrial and environmental applications. Required course for all students needing a first-year Chemistry course who have a grade of 67% or higher in both Chemistry 12 and Mathematics 12. OUC equivalent: CHEM 111. [3-3-0]

Lectures: 8:00am - 9:20am; Monday, Wednesday. Art 376
            8:30am - 9:20am; Friday. Art 376

Laboratory: Mandatory, schedule varies

Required Text / Material:
2. UBCO Chemistry 121/123 Lab Manual

Other Helpful Material:
1. Course web site: www.people.ok.ubc.ca/orcac/chem121.html
Chemistry 121: Topic 1 - Introduction

Evaluation:
- Quizzes 10%
- Laboratory 20%
- Midterms (x2) 30% (Oct. 6, Nov. 10)
- Final Exam 40% (Common Exam)

Note: The laboratory and the lecture must be passed independently.

It is your responsibility to review and understand the UBC policies on academic misconduct. The section describing the nature and consequences of academic misconduct are described in Chapter V of the UBC calendar:
http://okanagan.students.ubc.ca/calendar/index.cfm?tree=3,54,111,0

There is also page on the UBC web site describing this as well;
http://web.ubc.ca/okanagan/faculties/resources/academicintegrity.html

In addition to this UBC has some capacity to accommodate individuals with disabilities. In this context I have been requested to include the following comments.
"If you require disability related accommodations to meet the course objectives please contact the Coordinator of Disability Resources located in the Student development and Advising area of the student services building. For more information about Disability Resources or about academic accommodations please visit the website at www.okanagan.ubc.ca/current/disres.cfm"

To receive credit for Chemistry 121 - You must obtain at least 50% in both the lecture and lab sections of the course considered separately.
Course Outline:

**Topic 1: Introduction**
- 1.0 Introduction and review
- 1.1 Scientific method
- 1.2 Physical quantities: measurement, units, sig. figs., unit conversion
- 1.3 Chemistry and the nature of matter
- 1.4 Symbols: elements, compounds

**Topic 2: From Atoms to Stoichiometry**
- 2.0 Classification of materials: mixtures, elements, compounds
- 2.1 The weight laws.
- 2.2 Atoms, molecules and ions.
- 2.3 Atomic/molecular masses, isotopes, mass spectroscopy
- 2.4 The mole
- 2.5 Inorganic formulae and nomenclature
- 2.6 % Composition, combustion analysis
- 2.7 Empirical and molecular formulae.
- 2.8 Mole and mass relationships from balanced chemical equations.
- 2.9 Yields; % yields; limiting reagents.
- 2.10 Volumetric and gravimetric analyses
- 2.11 Analysis of mixtures.
Chemistry 121: Topic 1 - Introduction

Topic 3: Atomic Structure and Periodicity
3.1 Nature of light, elementary spectroscopy.
3.2 The quantum theory and the Bohr atom.
3.3 Quantum mechanics; the orbital concept.
3.4 Electron configurations of atoms
3.5 The periodic table: its historical development.
3.6 The periodic table: atomic structure & periodic trends.

Topic 4: Chemical Bonding
4.1 The ionic and covalent bond types & properties
4.2 Lewis Dot structures and the octet rule (and exceptions)
4.3 Molecular geometry: the VSEPR approach
4.4 Bonding: valence bond theory, hybridization and geometry
4.5 Bonding: molecular orbital theory.
4.6 Formal charges; resonance

Topic 5: The Gaseous State
5.1 Empirical approach to the gas laws, Boyle, Charles.
5.2 The kinetic-molecular theory of gases (qualitatively).
5.3 Ideal vs. real gases.

Topic 6: Solids and Liquids
6.1 Four types of crystalline solids. Their properties using bond types.
6.2 Solid, liquid and vapour equilibria; the phase diagram.
6.3 Solubility: saturation, unsaturation, supersaturation; concentration
6.4 Effects of temperature and pressure.
<table>
<thead>
<tr>
<th>Time</th>
<th>Monday</th>
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Topic 1 Introduction: Text: Chapters 1  (3 lectures)
1.0 Course Introduction (completed)
1.1 Chemistry and the nature of matter
1.2 The Scientific Method
1.3 Mathematical Review
1.4 Physical Quantities: units, unit conversions
1.5 Measurement: uncertainty, accuracy, precision, significant figures, errors
1.6 Problem Solving

- What is Chemistry?
WHAT IS CHEMISTRY?

- **CHEMISTRY** is the **SCIENTIFIC** study of the properties of **MATTER**, including changes in these properties and the accompanying **ENERGY** changes.

- **MATTER** is anything that occupies space and has mass

- **ENERGY** is the capacity to do work

- **THE SCIENTIFIC METHOD** is the process of observation, hypothesis and verification

- Why is Chemistry called the Central Science?

- Describe what you think is the Scientific Method?
The Scientific Method: A systematic Approach to Research

A) Define the Problem

B) Create Working Hypothesis  (Hypo = under, less: Thesis = a proposition)

C) Perform Experiment - Collect Data
   Qualitative - General observations about system
   Quantitative - Numbers collected by measurements of the system

D) New Hypothesis based on data: tentative explanation of data

E) Test Hypothesis with new experiment (true many times = Theory)

Theory: A unified principle that explains a body of facts and/or those laws based on them.

Law: A concise verbal or mathematical statement of a relationship between phenomena that is always the same under the same conditions

- Water jar experiments
NUMBERS ARE CENTRAL TO SCIENCE

YOU CANNOT DO CHEMISTRY WITHOUT SOME BASIC NUMERICAL SKILLS. IT IS NOT A SOFT OPTION. YOU MUST DEMONSTRATE THESE SKILLS BEFORE WE START DOING THE CHEMISTRY OR YOU ARE WASTING TIME.
1.3.1 Exponential or Scientific Notation

- \(6100000000000000 = 6.1 \times 10^{15}\)
- \(0.0000023 = 2.3 \times 10^{-6}\)
- \(6.023 \times 10^{23} = 602,300,000,000,000,000,000,000,000\)

Is \(54.3 \times 10^2\) correct scientific notation?
1.3.1 Exponential or Scientific Notation (Mathematical Functions)

Add 20,000 to 2000 = 22,000

Add $2 \times 10^4$ to $2 \times 10^3$

\[
\begin{align*}
20 \times 10^3 \\
+ 2 \times 10^3 \\
\hline
22 \times 10^3
\end{align*}
\]

$= 2.2 \times 10^4$

\[
\begin{align*}
2 \times 10^4 \\
x 2 \times 10^3 \\
\hline
4 \times 10^7
\end{align*}
\]

$= 20000 \times 2000 = 40000000$

Of course, the answer must always be the same!
(2 \times 10^2)^2 = 4 \times 10^4  \quad \text{and} \quad (2 \times 10^2)^3 = 8 \times 10^6

\sqrt{4 \times 10^4} = 2 \times 10^2 \quad \text{also stated as} \quad (4 \times 10^4)^{0.5} \quad \text{or} \quad (4 \times 10^4)^{1/2}

(8 \times 10^6)^{1/3} = 2 \times 10^2
1.3.2 Logarithms

- 1000 in Scientific notation is $1 \times 10^3$
- $\log(1000) = 3$ (exponent part of notation)

Therefore;

$log(100000) = 5$ and $log (0.000001) = ?$

$log(0.000001) = log(1 \times 10^{-6}) = -6$

$log(1) = 0$ and $log(10) = 1$
Recall:

- Non-exponential part always between 1 and 10
- $\log(AB) = \log(A) + \log(B)$
- $\log$ (Exponent part) = exponent (ie. 23 above)
- $\log$ (Non-exponential) from Figure 1.1

![Figure 1.1 Shortcut log table](image)
1.3 Mathematical Review

- \( \log (6.023 \times 10^{23}) = 23 + 0.7798 = 23.7798 \)
- \( \log (2.0 \times 10^{23}) = 23 + 0.30 = 23.30 \)
- \( \log (2.0 \times 10^{-10}) = -10 + 0.30 = -9.70 \)

Other Logs Functions:

- \( \log(A/B) = \log(A) - \log(B) \)
- \( 10^A \times 10^B = 10^{A+B} \)

- \( \frac{10^A}{10^B} = 10^{A-B} \)

Natural Logarithms:
- logs taken to the base of \( e \) instead of 10
- \( e \) is equal to 2.7183
- \( \ln(10) = 2.303 : e^{2.303} = 10 \)
1.3.3 Significant Figures

There are two types of numbers;

- Counting numbers and Measured numbers

Counting Numbers exact, ie. 31 students in Chem. 111
Measured numbers have units and significant figures

Significant figures reflect the magnitude of certainty (error)

If an object is said to be 24.5 inches long, then the certainty of the measurement is also evident. ie., between 24.4 and 24.6 inches. The value 24.5 has 3 significant digits.

How many significant figures are in;
(a) 23 inches          (b) 0.045 mm
(c) 60001 Liters      (d) 6 people
(e) 0.00100 miles     (f) 100 Km
1.3.3 Significant Figures (Functions)

Addition & Subtraction (answer same digits after decimal as least)

\[
\begin{align*}
89.332 & \quad + \quad 1.1 \quad - \quad 0.12 \\
\underline{} & \quad \underline{} \quad \underline{} \\
90.432 & \quad \text{round to 90.4} \quad 1.977 & \text{round to 1.98}
\end{align*}
\]

Number of digits after decimal implies accuracy (ie. if inches above)

Multiplication & Division (answer same digits # sig. figs. as least)

\[
\begin{align*}
2.8 & \times 4.5039 = 12.61092 \rightarrow \text{round to 13} \\
(6.85) & \div (112.04) = 0.0611388789 \rightarrow \text{round to 0.0611} \ (6.11 \times 10^{-2})
\end{align*}
\]

Problem:

What is the mass of eight objects that weigh 1.000 Kilograms each?

\[
8 \times 1.000 = 8.000 \rightarrow \text{the number of objects is exact!}
\]

Round 6.114575 x 10^{-2} to; 3 Significant Figures (6.11 x 10^{-2})

Sig. Figs. (6.115 x 10^{-2}) \quad 6 \text{ Sig. Figs} \ (6.11458 \times 10^{-2})
1.4 Physical Quantities

SI Units used in Chemistry and most sciences are based on a French System 
Système International d’Unités (basis of SI) these units are:

<table>
<thead>
<tr>
<th>Fundamental SI units</th>
<th>Quantity</th>
<th>Unit (symbol)</th>
<th>Definition</th>
</tr>
</thead>
<tbody>
<tr>
<td>Length</td>
<td></td>
<td>meter (m)</td>
<td>One meter is the distance light travels in a vacuum during $\frac{1}{299,792,458}$ of a second.</td>
</tr>
<tr>
<td>Mass</td>
<td></td>
<td>kilogram (kg)</td>
<td>One kilogram is the mass of the prototype kilogram kept at Sèvres, France.</td>
</tr>
<tr>
<td>Time</td>
<td></td>
<td>second (s)</td>
<td>One second is the duration of 9 192 631 770 periods of the radiation corresponding to a certain atomic transition of $^{133}\text{Cs}$.</td>
</tr>
<tr>
<td>Electric current</td>
<td></td>
<td>ampere (A)</td>
<td>One ampere of current produces a force of $2 \times 10^{-7}$ newtons per meter of length when maintained in two straight, parallel conductors of infinite length and negligible cross section, separated by one meter in a vacuum.</td>
</tr>
<tr>
<td>Temperature</td>
<td></td>
<td>kelvin (K)</td>
<td>Temperature is defined such that the triple point of water (at which solid, liquid, and gaseous water are in equilibrium) is 273.16 K, and the temperature of absolute zero is 0 K.</td>
</tr>
<tr>
<td>Luminous intensity</td>
<td></td>
<td>candela (cd)</td>
<td>Candela is a measure of luminous intensity visible to the human eye.</td>
</tr>
<tr>
<td>Amount of substance</td>
<td></td>
<td>mole (mol)</td>
<td>One mole is the number of particles equal to the number of atoms in exactly 0.012 kg of $^{12}\text{C}$ (approximately 6.022 136 7 $\times 10^{23}$).</td>
</tr>
<tr>
<td>Plane angle</td>
<td></td>
<td>radian (rad)</td>
<td>There are $2\pi$ radians in a circle.</td>
</tr>
<tr>
<td>Solid angle</td>
<td></td>
<td>steradian (sr)</td>
<td>There are $4\pi$ steradians in a sphere.</td>
</tr>
</tbody>
</table>
In SI System large and small numbers related through descriptive prefixes:

For Example:
The unit of length in the SI system is the meter. (about 3 ft)
1/100 of this unit is the centimeter, about the width of a large paper clip
1/1000 of this unit is the millimeter, about the width wire in a large paper clip

1 meter = 100 centimeter (1 x10^2 cm) = 1000 millimeter (1 x10^3 mm)

Common Prefixes (You Should Know):

<table>
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<tr>
<th>Prefix</th>
<th>Symbol</th>
<th>Exponent</th>
</tr>
</thead>
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<tr>
<td>giga</td>
<td>G</td>
<td>10^9</td>
</tr>
<tr>
<td>mega</td>
<td>M</td>
<td>10^6</td>
</tr>
<tr>
<td>kilo</td>
<td>k</td>
<td>10^3</td>
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<td>centi</td>
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<td>10^{-2}</td>
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<td>milli</td>
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<td>micro</td>
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<td>10^{-9}</td>
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<td>pico</td>
<td>p</td>
<td>10^{-12}</td>
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<tr>
<td>femto</td>
<td>f</td>
<td>10^{-15}</td>
</tr>
</tbody>
</table>
Mass and Weight:
SI units are based on **Mass** not **Weight**. Weight is “..the force gravity exerts on an object.”. In a lower gravitational field your weight would be less but your mass would be the same.

- The SI unit for mass is the kilogram (kg);
  - 1 kg = 1000 grams = $1 \times 10^3$ g
  - The pound (lb) is commonly used: 1 lb = 0.454 kg

Volume:
Volume is an SI derived unit related to the meter. The common unit for volume is the Liter (L). Key relations for volume are:

- $1 \text{ L} = 1000 \text{ mL} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$
- $1 \text{ dm}^3$ (one decimeter cubed) = $(1 \times 10^{-1} \text{ m})^3 = 1 \times 10^{-3} \text{ m}^3$
- $1 \text{ mL} = 1 \text{ cm}^3$

Density:
Density is the ratio between mass and volume;
density $(d) = \frac{\text{mass}}{\text{volume}}$

Technically the SI derived units are Kg/m$^3$ however in practice gram/mL (g/cm$^3$) is most often used. For gasses gram/Liter is often used.
**Temperature:**
There are 3 temperature scales currently in use. The Fahrenheit scale is still in common use in the US, although it is a very poorly designed system.

The Kelvin scale is the SI base for temperature. It is also called the absolute temperature scale. It’s units are identical in magnitude to the commonly used Celsius scale, thus both are commonly used in chemistry.

To convert:  
\[ ^\circ C = (^\circ F - 32) \times \left(\frac{5^\circ C}{9^\circ F}\right) \]
\[ K = (^\circ C + 273.15) \]
1.5 Measurement and error

Measurement is a key component in Science. However, with every measurement except for counting the measured value contains some type of inherent error.
Error Types:

Random Error:
- Error that occurs randomly and that normally cannot be avoided. There is an equal probability of a positive and a negative deviation from the measured value.
- For Example, random electrical noise

Systematic or Determinate Error:
- A type of error with both a definite magnitude and sign.
- Typical sources are Instrumental errors, Personal errors, and Method errors.
Accuracy:
- Indicates how close a measured value is to the true value.
- Effected by both random and systematic error.
- Determined from a comparison of average measured values with known values.

Precision
- A measure of the reproducibility of the method.
  - How close are the results obtained in the same way?
- Determined from a comparison of measured values with the average measured value.
- Standard deviation is an effective monitor of random error.

Accurate & Precise

Precise not Accurate

Accurate not Precise
Problem Solving (Factor-Label Method and Dimensional Analysis):

All measured numbers must have units to carry any meaning. The careful consideration of these units (the dimensions) can provide insight into problem solving.

For example;

You know you live 3.0 miles away from a mall. Can you determine how many feet you are from the mall, given there are 5280 feet in a mile?

\[ 3.0 \text{ mile} \times \frac{5280 \text{ feet}}{\text{mile}} = 15840 \text{ feet} = 1.6 \times 10^4 \text{ feet} \]

Problem Solving Recipe:
1. Write down the given number(s) with its units
2. Write a ratio with the given number in the denominator (at the bottom) and the unit sought in the numerator (on top)
3. Insert numbers into the ratio such that numerator and denominator are equal
4. Multiply Steps 1 and 3 together
Problem Solving:

**Example 2:** Five weeks ago you mailed a letter to a friend and are still waiting for a reply. How many days have you waited?

- **Step 1:** 5.0 weeks
- **Step 2:** days/weeks (days per week)
- **Step 3:** You know 7 days = 1 week \( \therefore 7 \text{ days/week} \)
- **Step 4:** 5.0 weeks \( \times \) 7 days/week = 35 days

**Example 3:** You have purchased a new car which gets 80 mile per gallon. Your gas tank has only 5 gallons of gas in it. You want to make a 500 mile trip. How far can you travel with the available gas?

\[
5 \text{ gallons } \times 80 \text{ mile/gallon} = 400 \text{ miles} = 4 \times 10^2 \text{ miles}
\]

- Notice all 4 steps can be combined into one step.
**Problem Solving:**

*Example 3:* Typical seawater contains 2.7 grams of salt (sodium chloride, NaCl) for every 100 ml. Given that the molecular weight of NaCl is 58.44 g/mole and that molarity is defined as the number of moles per liter; what is the molarity of NaCl in the ocean?

Know: \( \frac{2.7 \text{ grams}}{100 \text{ ml}} : \frac{58.44 \text{ grams}}{\text{mole}} : \frac{1000 \text{ ml}}{\text{Liter}} \)

Molarity = mole/Liter

How many grams per Liter?
\[
(2.7 \text{ grams} / 100 \text{ ml}) \times (1000 \text{ ml} / \text{Liter}) = 27 \text{ grams} / \text{Liter}
\]

How many moles in one Liter?
\[
27 \text{ grams} \div 58.44 \text{ grams/mole} = 0.46 \text{ moles}
\]

∴ \( 0.46 \text{ moles} / \text{Liter} = 0.46 \text{ Molarity of NaCl in the Ocean} \)