Chemistry 151

• Professor James H. Geiger
• Office: Chemistry Building, Room 9
• Office Hours: 1:30-2:30 PM MWF, and other times by appointment (send me an email).
• You can also drop by, but I might be busy.
• Email: geigerj@msu.edu
Textbooks/other help

• Textbooks
• The same text will be used for CEM 152 in the spring semester.
• The 9th edition (2003) can also be used.
• The 10th edition is stocked by campus bookstores. Also, it can be ordered from Amazon.com, barnesandnoble.com, or directly from the publisher.
• N.B.: The study guide, Student's Guide for Chemistry, the Central Science, 10th edition, Prentice-Hall, 2005, is also available at the campus bookstores, but purchase of this book is not required.
• Lecture notes will be available on the web.
Course organization

• Lectures MWF 12:40-1:30 pm (me)

Recitation once a week (check your schedule). Small class, more individual help from Teaching assistants. Each section = 1 recitation group.

No Recitation this week.
They start next week.
This week only come to class WF 12:40-1:30 pm.
Grades

• Four exams (150 points/exam)\times 4 = 600

• Many quizzes (in class), once a week at least. (Total = 200). At least the lowest two quizzes will be dropped.

• **There will be no makeups.**
  – quiz problems *will be directly copied from homework problems*, except the numerical values will be changed such that the numerical answer is different.

  Final exam (200 points). Will be given on exam week.
How to succeed:

- Attend lecture and recitation
- Do homework problems
- Do extra problems if you think you need them
- **Being able to do the problems is key**
- Understand the concepts from lecture.
Lectures

• Will follow the book closely
• Example problems will be a key part.
Topics to be covered

First 9 chapters, Chapter 24 and 25

Chap 1 matter and measurement
- Chap 2, Atoms, molecules and Ions
- Chap 3 Stoichiometry, The Mole!
- Chap 4, reactions in water and solution stoichiometry
- Chap 5, Thermochemistry
- Chap 6, Electronic structure, atoms
- Chap 7, The periodic table
- Chap 8, Chemical bonding
- Chap 9, Molecular geometry
- Chap 24, Coordination chemistry
- Chap 25, Organic and biological chemistry
Chapter 1
Introduction:
Matter and Measurement
Scientific Method:

A systematic approach to solving problems.

- **Empirical**
  - Observations and experiments
  - Find patterns, trends, and laws
  - Formulate and test hypothesis
  - Theory

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**The testing and retesting**

**This is what makes it Science!**
Chemistry:

The study of matter and the changes it undergoes.
Matter:

Anything that has mass and takes up space.
Atoms are the building blocks of matter.
• **Atoms** are the building blocks of matter.
• **Each element** is made of the same kind of atom.
Atoms are the building blocks of matter.
Each element is made of the same kind of atom.
A compound is made of two or more different kinds of elements.
States of Matter

Gas: Cool or compress, Heat or reduce pressure

Liquid: Cool, Heat

Crystalline solid: Heat
Classification of Matter

Matter

Is it uniform throughout?

Mud

Heterogeneous mixture

salt water

Homogeneous

Does it have a variable composition?

Substances

water

Pure substance

water

Can it be separated into simpler substances?

oxygen

Element

Compound

salt water

Homogeneous mixture (solution)
Mixtures and Compounds

Element
(atoms)

Element
(molecules)

Compound
(molecules)

Mixture

(a) Atoms of an element  
He, Ne

(b) Molecules of an element  
N\textsubscript{2}, O\textsubscript{2}, Cl\textsubscript{2}

(c) Molecules of a compound  
CO\textsubscript{2}, H\textsubscript{2}O, NH\textsubscript{3}

(d) Mixture of elements and a compound  
Mix

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Properties and Changes of Matter
Properties of Matter

• Physical Properties:
  □ Must be observed without changing a compound/element into another compound/element.
    • Boiling point, density, mass, volume, etc.

• Chemical Properties:
  □ Can *only* be observed when a compound/element is changed into another compound/element.
    • Flammability, corrosiveness, reactivity with acid, etc.
Properties of Matter

• Intensive Properties:
  □ Independent of the amount of the matter that is present.
    • Density, boiling point, color, etc.

• Extensive Properties:
  □ Dependent upon the amount of the matter present.
    • Mass, volume, energy, etc.
Changes of Matter

• Physical Changes:
  □ Changes in matter that do not change the composition of a substance.
    • Changes of state, temperature, volume, etc.

• Chemical Changes:
  □ Changes that result in new substances.
    • Combustion, oxidation, decomposition, etc.
In the course of a chemical reaction, the reacting substances are converted to new substances.
Compounds

Compounds can be broken down into elements.
Relative abundance of elements

(a) Earth’s crust
- Oxygen: 49.5%
- Silicon: 25.7%
- Aluminum: 7.5%
- Iron: 4.7%
- Calcium: 3.4%
- Other: 9.2%

(b) Human body
- Oxygen: 65%
- Carbon: 18%
- Hydrogen: 10%
- Other: 7%
### TABLE 1.1 The Top Ten Chemicals Produced by the Chemical Industry in 2002

<table>
<thead>
<tr>
<th>Rank</th>
<th>Chemical</th>
<th>Formula</th>
<th>2002 Production (billions of pounds)</th>
<th>Principal End Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>Sulphuric acid</td>
<td>H₂SO₄</td>
<td>81</td>
<td>Fertilizers, chemical manufacturing</td>
</tr>
<tr>
<td>2</td>
<td>Nitrogen</td>
<td>N₂</td>
<td>73</td>
<td>Fertilizers</td>
</tr>
<tr>
<td>3</td>
<td>Oxygen</td>
<td>O₂</td>
<td>53</td>
<td>Steel, welding</td>
</tr>
<tr>
<td>4</td>
<td>Ethylene</td>
<td>C₂H₄</td>
<td>52</td>
<td>Plastics, antifreeze</td>
</tr>
<tr>
<td>5</td>
<td>Lime</td>
<td>CaO</td>
<td>38</td>
<td>Paper, cement, steel</td>
</tr>
<tr>
<td>6</td>
<td>Propylene</td>
<td>C₃H₆</td>
<td>32</td>
<td>Plastics</td>
</tr>
<tr>
<td>7</td>
<td>Ammonia</td>
<td>NH₃</td>
<td>29</td>
<td>Fertilizers</td>
</tr>
<tr>
<td>8</td>
<td>Chlorine</td>
<td>Cl₂</td>
<td>25</td>
<td>Bleaches, plastics, water purification</td>
</tr>
<tr>
<td>9</td>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>24</td>
<td>Fertilizers</td>
</tr>
<tr>
<td>10</td>
<td>Sodium hydroxide</td>
<td>NaOH</td>
<td>20</td>
<td>Aluminum production, soap</td>
</tr>
</tbody>
</table>

*a Most data from *Chemical and Engineering News*, July 7, 2003, pp. 53, 56.

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**Acids**  
**Bases**  
**Pure elements**
Separation of Mixtures
Filtration:

Separates heterogeneous mixture, solid substances from liquids and solutions.
Distillation:

Separates homogeneous mixture of liquids on the basis of differences in boiling point.
Separates homogeneous mixtures on the basis of differences in solubility in a solvent, or in binding to a solid matrix.

Separation techniques were critical to the development of the basic theories of chemistry. How do we know there are homogeneous mixtures? We can separate them.
Units of Measurement
### SI Units

**Learn! symbols and all!**

<table>
<thead>
<tr>
<th>Physical Quantity</th>
<th>Name of Unit</th>
<th>Abbreviation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Mass</td>
<td>Kilogram</td>
<td>kg</td>
</tr>
<tr>
<td>Length</td>
<td>Meter</td>
<td>m</td>
</tr>
<tr>
<td>Time</td>
<td>Second</td>
<td>s&lt;sup&gt;a&lt;/sup&gt;</td>
</tr>
<tr>
<td>Temperature</td>
<td>Kelvin</td>
<td>K</td>
</tr>
<tr>
<td>Amount of substance</td>
<td>Mole</td>
<td>mol</td>
</tr>
<tr>
<td>Electric current</td>
<td>Ampere</td>
<td>A</td>
</tr>
<tr>
<td>Luminous intensity</td>
<td>Candela</td>
<td>cd</td>
</tr>
</tbody>
</table>

<sup>a</sup>The abbreviation sec is frequently used.

- **Système International d’Unités**
- Uses a different base unit for each quantity
# Metric System

Prefixes convert the base units into units that are appropriate for the item being measured.

**Learn! More important than it looks!!!**

<table>
<thead>
<tr>
<th>Prefix</th>
<th>Abbreviation</th>
<th>Meaning</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Giga</td>
<td>G</td>
<td>$10^9$</td>
<td>1 gigameter (Gm) = $1 \times 10^9$ m</td>
</tr>
<tr>
<td>Mega</td>
<td>M</td>
<td>$10^6$</td>
<td>1 megameter (Mm) = $1 \times 10^6$ m</td>
</tr>
<tr>
<td>Kilo</td>
<td>k</td>
<td>$10^3$</td>
<td>1 kilometer (km) = $1 \times 10^3$ m</td>
</tr>
<tr>
<td>Deci</td>
<td>d</td>
<td>$10^{-1}$</td>
<td>1 decimeter (dm) = 0.1 m</td>
</tr>
<tr>
<td>Centi</td>
<td>c</td>
<td>$10^{-2}$</td>
<td>1 centimeter (cm) = 0.01 m</td>
</tr>
<tr>
<td>Milli</td>
<td>m</td>
<td>$10^{-3}$</td>
<td>1 millimeter (mm) = 0.001 m</td>
</tr>
<tr>
<td>Micro</td>
<td>$\mu$(^a)</td>
<td>$10^{-6}$</td>
<td>1 micrometer ((\mu)m) = $1 \times 10^{-6}$ m</td>
</tr>
<tr>
<td>Nano</td>
<td>n</td>
<td>$10^{-9}$</td>
<td>1 nanometer (nm) = $1 \times 10^{-9}$ m</td>
</tr>
<tr>
<td>Pico</td>
<td>p</td>
<td>$10^{-12}$</td>
<td>1 picometer (pm) = $1 \times 10^{-12}$ m</td>
</tr>
<tr>
<td>Femtometer</td>
<td>f</td>
<td>$10^{-15}$</td>
<td>1 femtometer (fm) = $1 \times 10^{-15}$ m</td>
</tr>
</tbody>
</table>

\(^a\)This is the Greek letter mu (pronounced “mew”).
Volume

- The most commonly used metric units for volume are the liter (L) and the milliliter (mL).
  - A liter is a cube 1 dm (10 cm) long on each side.
  - A milliliter is a cube 1 cm long on each side.
Temperature:

Temperature is proportional to the average kinetic energy of the particles in a sample.

K.E. = \( \frac{1}{2}mv^2 \)
Temperature

- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - 0°C is the freezing point of water.
  - 100°C is the boiling point of water.
Temperature

- The Kelvin is the SI unit of temperature.
- It is based on the properties of gases.
- $0 \text{ K} = 0 \text{ K.E.}$
- There are no negative Kelvin temperatures.
- $K = ^\circ\text{C} + 273.15$
The Fahrenheit scale is not used in scientific measurements.

- °F = \(\frac{9}{5}(°C) + 32\)
- °C = \(\frac{5}{9}(°F) - 32\)
Density:

Physical property of a substance Intensive.

\[ d = \frac{m}{V} \]
# Density of selected substances

<table>
<thead>
<tr>
<th>Substance</th>
<th>Density (g/cm³)</th>
</tr>
</thead>
<tbody>
<tr>
<td>Air</td>
<td>0.001</td>
</tr>
<tr>
<td>Balsa wood</td>
<td>0.16</td>
</tr>
<tr>
<td>Ethanol</td>
<td>0.79</td>
</tr>
<tr>
<td>Water</td>
<td>1.00</td>
</tr>
<tr>
<td>Ethylene glycol</td>
<td>1.09</td>
</tr>
<tr>
<td>Table sugar</td>
<td>1.59</td>
</tr>
<tr>
<td>Table salt</td>
<td>2.16</td>
</tr>
<tr>
<td>Iron</td>
<td>7.9</td>
</tr>
<tr>
<td>Gold</td>
<td>19.32</td>
</tr>
</tbody>
</table>
Uncertainty in Measurement
Uncertainty in Measurements

Different measuring devices have different uses and different degrees of accuracy/precision.

Which are more accurate?
# Exact versus inexact numbers

<table>
<thead>
<tr>
<th>Exact</th>
<th>Inexact</th>
</tr>
</thead>
<tbody>
<tr>
<td>1000 g/kg</td>
<td>ruler</td>
</tr>
<tr>
<td>measure</td>
<td></td>
</tr>
<tr>
<td>2.54 cm/in</td>
<td>Temp. reading</td>
</tr>
<tr>
<td>12/dozen</td>
<td>volume or mass</td>
</tr>
<tr>
<td>any conversion</td>
<td>etc.</td>
</tr>
<tr>
<td>factor</td>
<td></td>
</tr>
</tbody>
</table>
Significant Figures

- The term *significant figures* refers to digits that were measured.
- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the precision of our answers.
Significant Figures

1. All nonzero digits are significant. (sig figs in red)
   423.444

2. Zeroes between two significant figures are themselves significant.
   42,300045  42,340.0025

3. Zeroes at the beginning of a number are never significant.
   00042345.0  0.00048

4. Zeroes at the end of a number are significant if a decimal point is written in the number.
   423000 versus: 423000. or: 423000.000
Significant Figures

• When addition or subtraction is performed, answers are rounded to the least significant decimal place.

\[
\begin{align*}
24.245 &+ 22.33488 = 46.580 \\
\end{align*}
\]

• When multiplication or division is performed, answers are rounded to the number of digits that corresponds to the least number of significant figures in any of the numbers used in the calculation.

\[
\begin{align*}
35.8750 \times 40.006800 &= 1435.24 \text{ (6 sig figs)} \\
1435.24395 &= 1435.24 \text{ (6 sig figs)}
\end{align*}
\]
Accuracy versus Precision

- **Accuracy** How close a measurement is to the true value. (How right you are)

- **Precision** How close measurements are to each other. (Reproducibility) Tends to lead to systematic errors.
Dimensional analysis

What do virtually all problems in chemistry have in common?

Convert centimeters to feet: 1 cm = ? feet
Know: 2.54 cm = 1 in, 12 in = 1 foot.

\[
\frac{1 \text{ in}}{2.54 \text{ cm}} \left( \frac{1 \text{ ft}}{12 \text{ in}} \right) = 0.032 \frac{\text{ft}}{\text{cm}}
\]
Dimensional Analysis

- What do I need on top?
- What do I need on the bottom?
- What do I know?
- How do I get there?
- Note: You will always be given the conversion factors you need, you don’t have to memorize them.
Dimensional analysis, examples

The speed of light is $2.998 \times 10^{10}$ cm/s. What is it in km/hr?

Know: 1 km = 1000m, 1 m = 100 cm, 60 min = 1 hr, 60 sec = 1 min

What do I need on top? kilometers

What do I need on the bottom? hours

$$2.998 \times 10^{10} \text{ cm/s} \left( \frac{1 \text{ m}}{100 \text{ cm}} \right) \left( \frac{1 \text{ km}}{1000 \text{ m}} \right) \left( \frac{60 \text{ sec}}{1 \text{ min}} \right) \left( \frac{60 \text{ min}}{1 \text{ hr}} \right) = 1.0892 \times 10^9$$
Dimensional analysis, examples

The Vehicle Assembly Building (VAB) at the Kennedy Space Center has a volume of: $3,666,500 \text{m}^3$. What is it in liters?

Know: $1 \text{ L} = 1 \text{ dm}^3, 1 \text{ dm} = 0.1 \text{ m}$

What do I need on top? \textit{Liters}

What do I need on the bottom? \textit{nothing}

$$3,666,500 \text{m}^3 = \left(\frac{\text{dm}}{0.1 \text{m}}\right)^3 \left(\frac{1 \text{L}}{1 \text{dm}^3}\right) = 3.6665 \times 10^9 \text{L}$$
Dimensional analysis, examples

An individual suffering from high cholesterol has 232 mg cholesterol per 100.0 mL of blood. How many grams of cholesterol in the blood, assuming a blood volume of 5.2 L?

Know: 1 L = 1000 mL, 1g = 1000mg

What do I need on top? grams

What do I need on the bottom? patient

\[
\frac{232 \text{ mg}}{100.0 \text{ mL}} \left( \frac{1000 \text{ mL}}{1 \text{ L}} \right) \left( \frac{5.2 \text{ L blood}}{\text{ patient}} \right) \left( \frac{1 \text{ g}}{1000 \text{ mg}} \right) = 12. \frac{g}{\text{ patient}}
\]