The Scientific Method
A plan of attack!

1. Making observations. Observations may be qualitative (the sky is blue; water is a liquid) or quantitative (water boils at 100°C; a certain chemistry book weighs 2 kilograms). A qualitative observation does not involve a number. A quantitative observation (called a measurement) involves both a number and a unit.
2. Formulating hypotheses. A hypothesis is a possible explanation for an observation.
3. Performing experiments. An experiment is carried out to test a hypothesis. This involves gathering new information that enables a scientist to decide whether or not the hypothesis is valid—that is, whether it is supported by the new information learned from the experiment. Experiments always produce new observations, and this brings the process back to the beginning again.

- **Theory** – hypotheses are assembled in an attempt at explaining “why” the “what” happened.
- **Model** – we use many models to explain natural phenomenon – when new evidence is found, the model changes!
- **Scientific Laws** – a summary of observed (measurable) behavior [a theory is an explanation of behavior]
  - Law of Conservation of Mass – mass reactants = mass products
  - Law of Conservation of Energy – (a.k.a. first law of thermodynamics) Energy CANNOT be created NOR destroyed; can only change forms.
  A law summarizes what happens; a theory (model) is an attempt to explain WHY it happens.
- **Precision and Accuracy**
  - **accuracy** – correctness; agreement of a measurement with the true value
  - **precision** – reproducibility; degree of agreement among several measurements.

Experimental error, when used in this context, has a very specific meaning and does not necessarily imply a mistake or blunder. If you know about a mistake or blunder, you can, at least in principle, fix the problem and eliminate the mistake. Some experimental error is intrinsic. While it can be minimized, it cannot be eliminated. A perfectly executed experiment with no mistakes or blunders, still has experimental error. Experimental error falls into two categories: determinate and indeterminate.

**Determinate errors** have a definite direction and magnitude and have an assignable cause (their cause can be determined). Determinate error is also called **systematic error**. Determinate error can (theoretically) be eliminated. Error will be either too high or too low.

**Indeterminate errors** arise from uncertainties in a measurement as discussed above. Indeterminate error is also called **random error**, or noise. Indeterminate error can be minimized but cannot be eliminated. Error will be both too high and too low.
To measure accuracy $\Rightarrow$ % error can be used

$$\text{% error} = \left( \frac{|\text{experimental results} - \text{accepted value}|}{\text{accepted value}} \right) \times 100$$

To measure precision $\Rightarrow$ standard deviation can be used

$$\text{standard deviation} = \sqrt{\frac{\sum \text{(deviations}^2)}{\# \text{ trials} - 1}}$$

$$\text{deviation} = |\text{mean} - \text{each trial}|$$

(there will be the same number of deviations as trials)

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**Error Analysis**

Jack determined the density of three different samples of the same liquid to be 12.0 g/mL, 11.9 g/mL, and 12.1 g/mL. Jill also determined the density of three different samples of the same liquid and obtained the following results: 10.4 g/mL, 9.9 g/mL and 10.2 g/mL.

a. The excepted value for the density of the liquid is 10.0 g/mL. Was Jack’s or Jill’s data more accurate?

b. Was Jack or Jill’s data more precise? What kind of error is indicated by Jack’s results.

c. Determine the standard deviation of Jill’s data.
Remember SIG FIGs
Evaluate the following with proper attention to significant figures.

a. $0.0143 \text{ kg} + 91.05 \text{ g} + 95 \text{ cg} + 243 \text{ mg} = \frac{3.21 \times 10^{-4} \times 4.32 \times 10^{-3} \times 6.5 \times 10^{6}}{4.11 \times 10^{-6}}$

b. 

\[ ^\circ F = ^\circ C \times 1.8 + 32 \quad K = ^\circ C + 273 \]

$\text{Density} = \frac{\text{mass}}{\text{volume}}$
Remember Dimensional Analysis & Common Conversions:
Fourteen carat gold is a mixture of gold and other metals, with gold present to the extent of 58.3%. If you have a 14-carat gold ring with a mass of 5.63 g, the mass of gold present is?

Classification of Matter
• Matter – takes up space, has mass, exhibits inertia
• States of Matter
  solid – rigid; definite shape and volume; molecules close together vibrating about fixed points! virtually incompressible
  liquid – definite volume but takes on the shape of the container; molecules still vibrate but also have rotational and translational motion and can slide past one another BUT are still close together! slightly compressible
  gas – no definite volume and takes on the shape of the container; molecules vibrate, rotate and translate and are independent of each other! VERY far apart! highly compressible
  vapor – the gas phase of a substance that is normally a solid or liquid at room
temperature
- fluid – that which can flow; gases and liquids
- Mixtures – can be physically separated
  - homogeneous – have visibly indistinguishable parts, solutions including air
  - heterogeneous – have visibly distinguishable parts
  - means of physical separation include: filtering, fractional crystallization, distillation, chromatography

- Pure substances – compounds like water, carbon dioxide etc. and elements. Compounds can be separated into elements by chemical means.

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MIXTURES
- Variable composition
- Components retain their characteristic properties
- May be separated into pure substances by physical methods
- Mixtures of different compositions may have widely different properties

HOMOGENEOUS MIXTURES
- Have same composition throughout
- Components are indistinguishable

HETEROGENEOUS MIXTURES
- Do not have same composition throughout
- Components are distinguishable

PURE SUBSTANCES
- Fixed composition
- Cannot be separated into simpler substances by physical methods
- Can only be changed in identity and properties by chemical methods
- Properties do not vary

COMPUNDS
- Can be decomposed into simpler substances by chemical changes, always in a definite ratio

ELEMENTS
- Cannot be decomposed into simpler substances by chemical changes

MATTER
Everything that has mass

Physical changes

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Dalton's Atomic Theory rests on the following postulates.

1. **All matter consists of tiny particles.** The existence of atoms was first suggested more than 2000 years before Dalton's birth. Atoms remained pure speculation through most of this time, although Newton used arguments based on atoms to explain the gas laws in 1687. (Newton's speculations about atoms in the *Principia* were carefully copied by hand into Dalton's notebooks.)

2. **Atoms are indestructible and unchangeable.** Atoms of an element cannot be created, destroyed, broken into smaller parts or transformed into atoms of another element. Dalton based this hypothesis on the law of conservation of mass and on centuries of experimental evidence. With the discovery of subatomic particles after Dalton's time, it

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Dalton's exceptional gift for recognizing and interpreting patterns in experimental data lead him from a problem in meteorology to the idea of atoms as fundamental constituents of matter. He realized the vital theoretical connection between atomic weights and weight relations in chemical reactions. He was the first to associate the ancient idea of atoms with stoichiometry.
became apparent that atoms could be broken into smaller parts. The discovery of nuclear processes showed that it was even possible to transform atoms from one element into atoms of another. But we don't consider processes that affect the nucleus to be chemical processes. The postulate is still useful in explaining the law of conservation of mass in chemistry. A slightly more restrictive wording is "Atoms cannot be created, destroyed, or transformed into other atoms in a chemical change".

3. **All atoms of the same element are identical. Elements are characterized by the mass of their atoms.** All atoms of the same element have identical weights, Dalton asserted. Atoms of different elements have different weights. With the discovery of isotopes, however, the postulate was amended to read, "Elements are characterized by their atomic number".

4. **Atoms combine in new ways during a chemical change. When elements react, their atoms combine in simple, whole-number ratios.** This postulate suggested a practical strategy for determining relative atomic weights from elemental percentages in compounds. Experimental atomic weights could then be used to explain the fixed mass percentages of elements in all compounds of those elements!

This effectively explained both the law of definite proportions and the law of multiple proportions. Some of the details of Dalton's original atomic theory are now known to be incorrect. But the core concepts of the theory (that chemical reactions can be explained by the union and separation of atoms, and that these atoms have characteristic properties) are foundations of modern physical science.

**J. J. Thomson’s Cathode Tube Experiment**

"cathode rays" pass from negative electrode towards positive electrode in an evacuated tube

- hypothesis: cathode rays are streams of electrons
- calculated mass to charge ratio for electrons by observing bending of cathode rays in electric and magnetic fields
- proposed the plum pudding model of the atom
How J. J. Thomson used properties of cathode rays to hypothesize properties of the electron.

<table>
<thead>
<tr>
<th>observations</th>
<th>hypothesis</th>
</tr>
</thead>
<tbody>
<tr>
<td>ray properties are independent of the cathode material</td>
<td>... cathode ray stuff is a component of all materials</td>
</tr>
<tr>
<td>cathode rays bend near magnets</td>
<td>... magnets bend the paths of moving charged particles; maybe cathode rays are streams of moving charged particles</td>
</tr>
<tr>
<td>rays bend towards a positively charged plate.</td>
<td>... cathode rays are streams of negative charges</td>
</tr>
<tr>
<td>rays impart a negative charge to objects they strike.</td>
<td></td>
</tr>
<tr>
<td>Cathode rays don't bend around small obstacles,</td>
<td>... cathode rays behave like streams of particles</td>
</tr>
<tr>
<td>cast sharp shadows,</td>
<td></td>
</tr>
<tr>
<td>can turn paddlewheels placed in their path, and travel in straight lines</td>
<td></td>
</tr>
<tr>
<td>Measured the Charge/Mass ratio</td>
<td></td>
</tr>
</tbody>
</table>

Probing Atomic Structure: Cathode Rays

![Diagram of cathode rays experiment]

In 1909 Robert Millikan (1868-1953), working at the university of Chicago, performed a clever experiment to determine the magnitude of the electron charge. With the value of the charge to mass ratio determined by J. J. Thomson of $1.76 \times 10^8 \text{ C/g}$ Millikan was able to calculate the mass of the electron as $9.11 \times 10^{-31} \text{ kg}$. Millikan used an atomizer to produce charged oil drops and charged plates to stop the fall of the electrons due to gravity. The voltage and the mass of the oil drops were used to calculate the charge on the oil drop. Millikan’s data showed that the charge on the oil drops is always a multiple of a whole number of the electron’s charge.
Ernest Rutherford is one of the most illustrious scientists of all time. He is to the atom what Darwin is to evolution, Newton to mechanics, Faraday to electricity and Einstein to relativity. He radically altered our understanding of nature on three separate occasions. Through brilliantly conceived experiments, and with special insight, he explained the perplexing problem of radioactivity as the spontaneous disintegration of atoms (they were not necessarily stable entities as had been assumed since the time of the ancient Greeks), he determined the structure of the atom and he was the world's first successful alchemist (he converted nitrogen into oxygen). Or put another way, he was first to split the atom.

Discovery of the Nucleus

**hypothetical description of alpha particles** based on properties of alpha radiation

<table>
<thead>
<tr>
<th>observation</th>
<th>hypothesis</th>
</tr>
</thead>
<tbody>
<tr>
<td>alpha rays don't diffract</td>
<td>... alpha radiation is a stream of particles</td>
</tr>
<tr>
<td>alpha rays deflect towards a negatively charged plate and away from a positively charged plate</td>
<td>... alpha particles have a positive charge</td>
</tr>
<tr>
<td>alpha rays are deflected only slightly by an electric field; a cathode ray passing through the same field is deflected strongly</td>
<td>... alpha particles either have much lower charge or much greater mass than electrons</td>
</tr>
</tbody>
</table>
Scattering Experiment

- hypothesis: If the plum pudding model of the atom is correct, atoms have no concentration of mass or charge (atoms are 'soft' targets)
- experiment to test hypothesis:
  - fire massive alpha particles at the atoms in thin metal foil
  - alpha particles should pass like bullets straight through soft plum pudding atoms
- observation: a few alpha particles ricocheted!
- new hypotheses:
  - all of the positive charge and nearly all of the mass of the atom is concentrated in a tiny, incredibly dense 'nucleus', about $10^{-14}$ m in diameter electrons roam empty space about $10^{-10}$ m across, around the nucleus

The Nuclear Atom

- Thomson's Atom
  - diffuse mass and charge

- Rutherford's Atom
  - concentrated mass and positive charge at the nucleus
  - electrons roam empty space around the nucleus
• Composition of the Nucleus
  o nuclei are composed of "nucleons": protons and neutrons
  o atomic mass units
    • 1 amu (aka 1 dalton) = exactly 1/12 the mass of a carbon-12 nucleus
    • 1 amu = $1.67 \times 10^{-24}$ g

<table>
<thead>
<tr>
<th>Particle</th>
<th>Mass</th>
<th>Charge</th>
</tr>
</thead>
<tbody>
<tr>
<td>e^-</td>
<td>$9.11 \times 10^{-31}$</td>
<td>1-</td>
</tr>
<tr>
<td>p^+</td>
<td>$1.67 \times 10^{-27}$</td>
<td>1+</td>
</tr>
<tr>
<td>n^-</td>
<td>$1.67 \times 10^{-27}$</td>
<td>None</td>
</tr>
</tbody>
</table>

- **Mass Number**
  \[ ^A_Z\text{E} \]  
  Symbol for the element

- **Atomic Number**
  \[ ^{\text{atomic number}}_Z\text{X} \]  
  \[ Z \]  

- **atomic number**\((Z)\)--The number of p+ in an atom. All atoms of the same element have the same number of p+.
- **mass number**\((A)\)--The sum of the number of neutrons and p+ for an atom. A different mass number does not mean a different element--just an isotope.
• **average atomic masses**—atoms have masses of whole numbers, HOWEVER samples of quadrillions of atoms have a few that are heavier or lighter [isotopes] due to different numbers of neutrons present

• **percent abundance**—percentage of atoms in a natural sample of the pure element represented by a particular isotope

\[
\text{% abundance} = \left( \frac{\text{number of atoms of a given isotope}}{\text{total number of atoms of all the isotopes of that element}} \right) \times 100
\]

• **counting by mass**—when particles are small this is a matter of convenience. Just as you buy 5 lbs of sugar rather than a number of sugar crystals, or a pound of peanuts rather than counting the individual peanuts….this concept works very well if your know an average mass.

• **mass spectrometer to determine isotopic composition**—load in a pure sample of natural neon or other substance. The areas of the “peaks” or heights of the bars indicate the relative abundances of \( ^{20}\text{Ne} \), \( ^{21}\text{Ne} \), and \( ^{22}\text{Ne} \).

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### The Average Mass of an Element

When a sample of copper is vaporized and injected into a mass spectrometer, the results shown in the figure are obtained. Use this data to compute the average mass of copper. (The mass values for \( ^{63}\text{Cu} \) and \( ^{65}\text{Cu} \) are 62.93 amu and 64.93 amu, respectively.)
THE MOLE
- **mole**—the number of C atoms in exactly 12.0 grams of $^{12}$C; also a number, 6.02 x $10^{23}$ just as the word “dozen” means 12 and “couple” means 2.
- **Avogadro’s number**—6.02 x $10^{23}$, the number of particles in a mole of anything

MOLAR MASS, MOLECULAR WEIGHT, AND FORMULA WEIGHT
- **molar mass, MM**—the mass in grams of Avogadro’s number of molecules; i.e. the mass of a mole!
- **empirical formula**—that ratio in the network for an ionic substance.

### Calculating the Number of Moles and Mass

Cobalt (Co) is a metal that is added to steel to improve its resistance to corrosion. Calculate both the number of moles in a sample of cobalt containing 5.00 x $10^{20}$ atoms and the mass of the sample.

### Calculating Molar Mass I

Juglone, a dye known for centuries, is produced from the husks of black walnuts. It is also a natural herbicide (weed killer) that kills off competitive plants around the black walnut tree but does not affect grass and other noncompetitive plants [a concept called *allelopathy*]. The formula for juglone is C$_{10}$H$_6$O$_3$.

- a. Calculate the molar mass of juglone.
- b. A sample of 1.56 x $10^{-3}$ g of pure juglone was extracted from black walnut husks. How many moles of juglone does this sample represent?

### Calculating Molar Mass II

Calcium carbonate (CaCO$_3$), also called *calcite*, is the principal mineral found in limestone, marble, chalk, pearls, and the shells of marine animals such as clams.

- a. Calculate the molar mass of calcium carbonate.
- b. A certain sample of calcium carbonate contains 4.86 moles. What is the mass in grams of this sample? What is the mass of the CO$_3^{2-}$ ions present?

### Molar Mass and Numbers of Molecules

Isopentyl acetate (C$_7$H$_{14}$O$_2$), the compound responsible for the scent of bananas, can be produced commercially. Interestingly, bees release about 1µg (1 x $10^{-6}$ g) of this compound when they sting. The resulting scent attracts other bees to join the attack. How many molecules of isopentyl acetate are released in a typical bee sting?

How many atoms of carbon are present?
ELEMENTS THAT EXIST AS MOLECULES
Pure hydrogen, nitrogen, oxygen and the halogens [the “gens” collectively—easier to remember!] exist as diatomic molecules under normal conditions. MEMORIZE!!! Be sure you compute their molar masses as diatomics. Others to be aware of, but not memorize:
- P₄—tetratomic form of elemental phosphorous
- S₈—sulfur= elemental form
- Carbon—diamond and graphite networks of atoms

PERCENT COMPOSITION OF COMPOUNDS

Calculating Mass Percent I
Carvone is a substance that occurs in two forms having different arrangements of the atoms but the same molecular formula (C₁₀H₁₄O) and mass. One type of carvone gives caraway seeds their characteristic smell, and the other type is responsible for the smell of spearmint oil. Compute the mass percent of each element in carvone.

DETERMINING THE FORMULA OF A COMPOUND

When faced with a compound of “unknown” formula, one of the most common techniques is to combust it with oxygen to produce CO₂, H₂O, and N₂ which are then collected and weighed.

- **empirical and molecular formulas**: assume a 100 gram sample if given % of
  - empirical gives smallest ratio
  - need to know molar mass to establish molecular formula which is (empirical formula)ₙ, where n is an integer

- **determining empirical and molecular formulas**
  - *hydrates*—A dot waters
  - *anhydrous*—without water
A compound is composed of carbon, nitrogen and hydrogen. When 0.1156 g of this compound is reacted with oxygen [burned, combusted], 0.1638 g of carbon dioxide and 0.1676 g of water are collected. What is the empirical formula of the compound?

One last trick of the trade: When you don’t know the mass of your sample, assume 100 grams so that any percents become grams….proceed by finding the number of moles!

**Empirical Formula Determination**

- Since mass percentage gives the number of grams of a particular element per 100 grams of compound, base the calculation on 100 grams of compound. Each percent will then represent the mass in grams of that element.
- Determine the number of moles of each element present in 100 grams of compound using the atomic masses of the elements present.
- Divide each value of the number of moles by the smallest of the values. If each resulting number is a whole number (after appropriate rounding), these numbers represent the subscripts of the elements in the empirical formula.
- If the numbers obtained in the previous step are not whole numbers, multiply each number by an integer so that the results are all whole numbers.

**Determining Empirical and Molecular Formulas**

A white powder is analyzed and found to contain 43.64% phosphorus by mass and the rest being oxygen. The compound has a molar mass of 283.88 g/mol. What are the compound’s empirical and molecular formulas?
BALANCING CHEMICAL EQUATIONS

\[ \text{CH}_4 + \text{O}_2 \rightarrow \text{H}_2\text{O} + \text{CO}_2 \]

**Table 3.2 Information Conveyed by the Balanced Equation for the Combustion of Methane**

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>( \text{CH}_4(g) + 2\text{O}_2(g) )</td>
<td>( \text{CO}_2(g) + 2\text{H}_2\text{O}(g) )</td>
</tr>
<tr>
<td>1 molecule ( \text{CH}_4 )</td>
<td>1 molecule ( \text{CO}_2 ) + 2 molecules ( \text{H}_2\text{O} )</td>
</tr>
<tr>
<td>+ 2 molecules ( \text{O}_2 )</td>
<td></td>
</tr>
<tr>
<td>1 mol ( \text{CH}_4 ) molecules + 2 mol ( \text{O}_2 ) molecules</td>
<td>1 mol ( \text{CO}_2 ) molecules + 2 mol ( \text{H}_2\text{O} ) molecules</td>
</tr>
<tr>
<td>( 6.022 \times 10^{23} \text{ CH}_4 ) molecules + ( 2(6.022 \times 10^{23}) \text{ O}_2 ) molecules</td>
<td>( 6.022 \times 10^{23} \text{ CO}_2 ) molecules + ( 2(6.022 \times 10^{23}) \text{ H}_2\text{O} ) molecules</td>
</tr>
<tr>
<td>16 g ( \text{CH}_4 ) + 2(32 g) ( \text{O}_2 )</td>
<td>44 g ( \text{CO}_2 ) + 2(18 g) ( \text{H}_2\text{O} )</td>
</tr>
<tr>
<td>80 g reactants</td>
<td>80 g products</td>
</tr>
</tbody>
</table>

**Balancing a Chemical Equation I**

Chromium compounds exhibit a variety of bright colors. When solid ammonium dichromate, \((\text{NH}_4)\text{Cr}_2\text{O}_7\), a vivid orange compound, is ignited, a spectacular reaction occurs, as shown in the two photographs on page 105. Although the reaction is actually somewhat more complex, let’s assume here that the products are solid chromium(III) oxide, nitrogen gas (consisting of \(\text{N}_2\) molecules), and water vapor. Balance the equation for this reaction.

At 1000°C, ammonia gas, \(\text{NH}_3(g)\), reacts with oxygen gas to form gaseous nitric oxide, \(\text{NO}(g)\), and water vapor. This reaction is the first step in the commercial production of nitric acid by the Ostwald process. Balance the equation for this reaction.
Stoichiometry: Dimensional Analysis Method

A quantitative study of chemical changes.

The KEY to Stoichiometry is using the coefficients in the balanced chemical equation in the unit moles to convert from moles of one substance to moles of another substance.

**BASED UPON MOLE - MOLE RATIOS.**

**Basic Pattern:**

1. If quantity given is grams Convert to moles
2. Perform mole $\leftrightarrow$ mole conversion
3. Convert to desired units

If iron pyrite, FeS$_2$, is not removed from coal, oxygen from the air will combine with both the iron and the sulfur as coal burns. If a furnace burns an amount of coal containing 100 g of FeS$_2$, how much SO$_2$ (an air pollutant) is produced?

How many grams of oxygen are produced when 10.0 g of potassium chlorate are completely decomposed?

**LIMITING REAGENT PROBLEMS**

Is the reagent that is totally consumed during the reaction and therefore determines or limits the amount of product formed.

**To determine which reactant is the limiting reagent:**

1. Determine moles of each reactant
2. Compare the number of moles of each reactant to the ratio in which they will react to determine which will be totally consumed. Divided the number of moles of each reactant by the reactant’s coefficient in the balanced equation.
3. The smallest number refers to the limiting reagent.

The theoretical yield is the maximum amount of product which can be produced (in an ideal world). In the "real" world it is difficult to produce the amount obtained for the theoretical yield. A percent yield is often used to show how close to ideality one has obtained in a chemical synthesis.

A 50.6 g sample of Mg(OH)$_2$ is reacted with 45.0 g of HCl according to the reaction: Mg(OH)$_2 + 2$ HCl $\rightarrow$ MgCl$_2 + 2$ H$_2$O what is the theoretical yield of MgCl$_2$?

Is this a limiting reagent problem? One way to find out is to write down what is known about any component of the reaction below that component:
Again, you should not jump to any conclusions about which reactant is the limiting reagent. Just because there are fewer moles of magnesium hydroxide does not mean it is the limiting reagent.

PERCENT YIELD PROBLEMS:
- **Theoretical yield** – this is what you calculate from stoichiometry only in a perfect world
- **Actual yield** – what you get in a lab in the real world *(remember, energy is always lost and no process is 100% efficient, therefore no yield will be 100% useful in lab reports)*
- **Percent yield** = \[
  \frac{\text{Actual Yield (measured amount)}}{\text{Theoretical Yield (calculated value)}} \times 100
\]

Suppose in the reaction discussed a chemist actually obtained 55.4 g of MgCl₂. This is called the actual yield and would be given to you in the problem.

**Amount of Excess Reagent Remaining Unreacted:**
To determine the amount of excess reagent that remains un-reacted:

\[
\text{INITIAL AMOUNT} - \text{EXCESS AMOUNT} = \text{AMOUNT UNREACTED}
\]

Calculate the amount consumed using the limiting reagent.

\[
45.0 \text{ g HCl} = \frac{1 \text{ mole HCl}}{36.5 \text{ g HCl}} \times \frac{1 \text{ mole Mg(OH)}_2}{2 \text{ mole HCl}} \times \frac{58.3 \text{ g Mg(OH)}_2}{1 \text{ mole Mg(OH)}_2} = 35.9 \text{ g Mg(OH)}_2
\]

\[
\text{INITIAL AMOUNT} - \text{EXCESS AMOUNT} = \text{AMOUNT UNREACTED}
\]

\[
50.6 \text{ g} - 35.9 \text{ g} = 14.7 \text{ g}
\]

The space shuttle uses aluminum metal and ammonium perchlorate in its reusable booster rockets. The products of the reaction are aluminum oxide, aluminum chloride, nitrogen monoxide, and steam. The reaction mixture contains 5.75g aluminum and 7.32g of ammonium perchlorate.

a. Write the balanced chemical equation.

b. Identify the limiting reagent and the excess reagent. Then determine how many grams of aluminum chloride are formed.

c. Determine the number of grams of excess reagent remaining after the reaction.

d. If 1.87g of aluminum chloride is actually collected, what is the percent yield?
The Exxon Valdez spilled about $4.24 \times 10^{12}$ liters of oil. Exxon always gave the size of the spill in barrels, how many barrels is this? A barrel of oil contains 42 gallons.
Make the following measurement.
If gold has a density of 19.3g/mL how many atoms would there be in a sample of gold with a volume of .58 in³?
A particular type of brass contains the elements copper, tin, lead and zinc. A sample weighing 1.713 grams is treated in such a way as to convert the tin to 0.245 g of tin (IV) oxide, the lead to 0.115 g of lead (II) sulfate, and the zinc to 0.246 g of Zn$_2$P$_2$O$_7$. What is the percent composition of the brass?
A compound methyl butanoate smells like apples. Its percent composition is 58.8% C, 9.8 % H, and 31.4% O. Its molar mass is 102. g/mole. What is its molecular formula
A 0.151 g sample of a compound of carbon and hydrogen produces 0.5008 grams of carbon dioxide on combustion. If the molar mass is 106 g/mole, what is the molecular formula
A hydrocarbon mixture consists of 60.0% C₃H₈ and 40% CₓHᵧ. When 10.00 grams of this mixture is burned 29.0 g of carbon dioxide and 18.8 grams of water were produced. What is the empirical formula of CₓHᵧ?
A common type of fire extinguisher depends upon the reaction of sodium hydrogen carbonate, NaHCO₃, with sulfuric acid to produce carbon dioxide that develops a pressure to squirt water or foam onto a fire. The equation is

\[ 2 \text{NaHCO}_3 + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 2 \text{H}_2\text{O} + 2 \text{CO}_2. \]

If a fire extinguisher were designed to hold 600 grams of sodium hydrogen carbonate, how many grams of sulfuric would be required to react with all of it?
Copper is extracted from chalcocite, a copper(I) sulfide ore, by a reaction with oxygen that may be represented by the equation \[ \text{Cu}_2\text{S} + \text{O}_2 \rightarrow 2 \text{Cu} + \text{SO}_2 \]. If treatment of 41.9 grams of pure \text{Cu}_2\text{S} by the process yields 29.2 grams of copper, calculate the percentage yield.
Carbon disulfide, \( \text{CS}_2 \), burns in oxygen. Complete combustion gives the reaction \( \text{CS}_2(g) + 3\text{O}_2(g) \rightarrow \text{CO}_2(g) + 2\text{SO}_2(g) \). Calculate the grams of sulfur dioxide, \( \text{SO}_2 \), produced when a mixture of 15.0 g of carbon disulfide and 35.0 g of oxygen reacts. Which reactant remains unconsumed at the end of the combustion? How many grams remain?
To find the formula of a compound composed of iron and carbon monoxide, \( \text{Fe}_x(\text{CO})_y \), you burn the compound in pure oxygen according to the following, \textit{unbalanced} equation.

\[
\text{Fe}_x(\text{CO})_y(s) + \text{O}_2(g) \rightarrow \text{Fe}_2\text{O}_3(s) + \text{CO}_2(g)
\]

If you burn 1.959 g of \( \text{Fe}_x(\text{CO})_y \) and find 0.860 g of \( \text{Fe}_2\text{O}_3 \) and 2.133 g of \( \text{CO}_2 \), what is the empirical formula of \( \text{Fe}_x(\text{CO})_y \)?
Hydrated salts are very common. If you heat 2.105 g of $\text{CoCl}_2 \cdot x \text{H}_2\text{O}$, and find that 1.149 g of $\text{CoCl}_2$ remains, what is the value of $x$?