HONORS CHEMISTRY
MID-TERM REVIEW OUTLINE, (2010 – 2011)

Topic

1. Measurement in chemistry

   a) accuracy & precision

   - Accuracy - measure of how closely your data agrees with the true value.
     - Percent Error measures accuracy.
       - Formula:
       - Percent Error = $\frac{|True - Experimental|}{True} \bullet 100$
   - Precision – how closely your data agrees with itself
   - All measurements have a degree of uncertainty regardless of precision and accuracy such as the measuring instrument and the skill of the person who made the measurement.

   b) SI system – **(temp conversions- just Celsius and Kelvin)**

   - Same units used worldwide. Also known as the metric system.
   - It is so widely accepted because it uses the decimal system as its base and many units for various quantities are defined in terms of units for simpler quantities.

SI Base Units/ 7 Fundamental Units

Property	Unit	Abbreviation
mass	kilogram	kg
length	meter	m
time	second	s
electric current	ampere	A
temperature	Kelvin	K
amount of substance	mole	mol
luminous intensity	candela	cd

Aren't as important

- Other SI units are derived by combining prefixes with a base unit. The prefixes represent multiples or fractions of 10. This includes kilometer(km), milliliter (mL), and milligram (mg).
- Kilo means to multiply the root by 1,000, so a kilometer is 1,000 meters.

- **Some metric equivalents**
    - **1 g = 1 mL (so 1g=1cm^3)**
    - **1 mL = 1 cm^3 or 1 L = 1 dm^3**
    - 2.54 cm = 1 in

SI Prefixes

Know from tera to pico including the symbol and exponent (factor).

Factor	Name	Symbol
10^{12}	tera	T
10^{9}	giga	G
10^{6}	mega	M
10^{3}	kilo	k
10^{2}	hecto	h
10^{1}	deka	da
10^{-1}	deci	d
10^{-2}	centi	c
10^{-3}	milli	m
10^{-6}	micro	μ
10^{-9}	nano	n
10^{-12}	pico	p

Think of computer storage/memory from large to small – **<u>T</u>**era **<u>G</u>**iga **<u>M</u>**ega (all after this have lowercase symbols)

<u>K</u>ing **<u>H</u>**enry **<u>D</u>**ied By **<u>D</u>**rinking **<u>C</u>**hocolate **<u>M</u>**ilk

Last: **<u>M</u>**icro(μ) **<u>N</u>**ano **<u>P</u>**ico

<u>Mnemonic to Remember Prefixes</u>

<u>T</u>he **<u>G</u>**reat **<u>M</u>**an **<u>K</u>**ing **<u>H</u>**enry **<u>D</u>**ied By **<u>D</u>**rinking **<u>C</u>**hocolate **<u>M</u>**ilk; **<u>M</u>**ight **<u>N</u>**eed **<u>P</u>**rince.

- The First and Last 3 Factors of the Prefixes decrease by 3. The Middle 6 Factors of the Prefixes decrease by 1.
  - **Factors**:

12, 9, 6	3, 2, 1, -1, -2, -3	-6, -9, -12
3's	*1's*	*3's*

- Temperature Conversions – Kelvin and Celsius
  - K = ^{O}C + 273^{O}
  - ^{O}C = K - 273^{O}
  - *The freezing point is 0^{O}C and 273 K, boiling 100^{O}*C and 373 K
- When adding in scientific notation, the exponents must be the same. **Also moving the decimal point to the left is positive(increasing) and to the right is negative(decreasing).**

c) significant figures ✓

- Any time a measurement is recorded, it includes all the digits that are certain plus one uncertain digit. These certain digits plus the one uncertain digit are referred to as **significant figures.** The more digits you are able to record in a measurement, the less relative uncertainty there is in the measurement.

- Significant Figures Rules

Rule	Example	Number of Sig Figs
All digits other than zeros are significant.	25 g 5.471 g	2 4
Zeros between nonzero digits are significant	309 g 40.06 g	3 4
Final zeros to the right of the decimal point are significant.	6.00 mL 2.350 mL	3 4
In #'s smaller than 1, zeros to left or directly to the right of the decimal point are not significant.	0.05 cm 0.060	1 – the zeros merely mark the position of the decimal point 2 – The first two zeros mark the position of the decimal point. The final zero is significant.

- Calculations with significant figures – when doing calculations with numbers that do not have the same # of sig figs, keep these rules in mind.
    - **<u>Multiplication and Division</u>** – the number of sig figs in a product or quotient of measured quantities is the same as the # of sig figs in the quantity having the smaller # of sig figs
        - **<u>Ex.</u> – 4.29 cm x 3.2 cm = 13.728 (unrounded) = 14 cm^2 (2 sig figs)**
    - **<u>Addition and Subtraction</u> – should be rounded to the same number of decimal places as the quantity having the least number of decimal places.**
        - **<u>Ex.</u>**

$$\begin{array}{r} 3.56 \text{ cm} \\ 2.6 \text{ cm} \\ +6.12 \text{ cm} \\ \hline \end{array}$$

12.28 (unrounded) = 12.3 cm (1 decimal place)

2. Unit analysis-problem solving ✓

- **Dimensional Analysis** – Changing from one unit to another via conversion factors

(based on the equivalence statements between the units)

- Used to solve problems (way to solve them)
  - Ex. 25,451 cm -----------> km

1 km - __10^5__ cm

km cm
10^3 10^{-2} = 10^5

Need

1. Conversion factor – relationship between 2 units
2. Given information

$$25{,}451 \text{ cm} \times \frac{1 \text{ km}}{10^5 \text{ cm}} = 25{,}451 \times 10^{-5} \text{ km}$$

$\boxed{2.5451 \times 10^{-1} \text{ km}}$

Not in Scientific Notation

Final Answer, In Scientific Notation

- **When you are doing dimensional analysis it is 1 of the bigger number equals how many of the smaller number (ex. 1 m = 10^2 cm).**

3. Density ✓

- **Density = Density** $= \frac{\textbf{mass}}{\textbf{volume}}$ **or D** $= \frac{\textbf{m}}{\textbf{v}}$
- The basic unit of mass is grams and for volume is cm^3 or mL. **Density is usually g/cm^3**
- **Defined as mass per unit volume**.

4. Classification of matter - **Topic 2-I**

a) states & changes of matter-S,L,G

- **States of matter** – solid, liquid, gas. Change of state is a physical change.

sublimation **deposition**
SOLID ============> GAS ==============> SOLID

melting **freezing**
SOLID ============> LIQUID =============> SOLID
(solidification)

condensation **vaporization**
GAS =============> LIQUID =============> GAS

- o Sublimation, deposition
  - ▪ **Sublimation –** change of state straight from solid to gas
    - Ex. dry ice, solid air freshener
  - ▪ **Deposition –** change of state straight from gas to solid
- o Boiling point, Freezing point, Melting, Condensation
  - ▪ **Boiling point** - the temperature at which a liquid boils and turns to vapor. (A change of state from liquid to gas by adding heat/energy absorbed called vaporizing. Example boiling water.)
  - ▪ **Freezing point** - the temperature at which a liquid turns into a solid when cooled. (A change of state from liquid to solid by releasing heat/energy called freezing. An example is water to ice.)
  - ▪ **Melting** – change from solid to liquid by adding heat (energy absorbed). An example is ice to liquid water.
  - ▪ **Condensation** – change from gas to liquid as heat/energy is released. An example is rain.

b) elements, compounds, mixtures (homogeneous and heterogeneous), solutions (properties, definitions, separation) - p. 41
a solution is the same as a homogenous mixture

- **<u>Matter</u>**
  - o <u>Mixture</u> - **A substance composed of two or more components, each of which retains its own properties (can vary in their composition) -** homogeneous and heterogeneous mixtures
    - ▪ **Homogeneous (solution)** – uniform composition throughout
      - Examples
        - o Salt H_20

- H_2O_2 – Hydrogen Peroxide
  - Molecular compound
- Alcohol diluted with water
- Sterling silver (Ag Cu)

- **Heterogeneous** – not uniform composition throughout
  - Examples
    - Oil and Water (H_2O)
    - Salad
    - Dirt/sand in water
  - **The number of phases in a mixture**
    - Heterogeneous – two phases or more
    - Homogeneous – uniform, so always one phase

- **Pure substances** (elements and compounds, compounds being both molecular and ionic or salts)
  - **<u>Element</u> - One of the more than 100 "building blocks" of which all matter is composed. An element consists of atoms of only one kind and cannot be decomposed further by ordinary chemical means.** The word **atom** comes from the Greeks and means the smallest possible piece of something.

  - **<u>Compounds</u>** – **when two or more atoms of different elements are chemically combined in definite proportions by weight.**
    - The smallest naturally occurring unit of a compound is called a **molecule** of that compound. A molecule of a compound has a definite shape that is determined by how the atoms are bonded to or combined with each other. An example is the compound water: it always occurs in a two hydrogen atoms to one hydrogen atom relationship.
      - **Molecular** – atoms linked together by sharing electrons. Basically they bind together in electrically neutral particles called molecules. Some molecular compounds are very simple. The very examples of these are diatomic molecules, which only consists of two atoms. Carbon Monoxide (CO) is an example of a diatomic compound.
      - **Ionic** – a compound that results when a metal reacts with an nonmetal (to form cations and anions)
        - Salts – ionic compound example

- **<u>Solutions</u>**
    - Properties – see above
    - Definitions
        - **Solution – a uniform mixture of solute and solvent**
    - Separation – see below in chart on #5. **A solution is the same thing as a homogenous mixture.**

c) atoms, molecules, ions

- **Atom -** comes from the Greeks and means the **smallest possible piece of something**. Today, scientists recognize approximately 109 different kinds of atoms, each with its own unique composition. These atoms then are the building blocks of elements when only one kind of atom makes up the substance.
- **Molecule** - **The smallest naturally occurring unit of a compound is called a <u>molecule</u> of that compound**. A molecule of a compound has a definite shape that is determined by how the atoms are bonded to or combined with each other. An example is the compound water: it always occurs in a two hydrogen atoms to one hydrogen atom relationship.
    - Compounds -
        - **Molecular** – atoms linked together by sharing electrons.
        - **Ionic** – a compound that results when a metal reacts with an nonmetal (to form cations and anions)

- **Ion** – **an atom or group of combined atoms that carries one or more electric charges**. Examples NH_4^{1+},OH^{1-} . An ionic compound results when a metal reacts with a nonmetal (to form cations and anions). An example is salts.

d) classes of elements; symbols

- Metals, Nonmetals, Metalloids (also called semimetals)
    - **Metals** – an element that gives up electrons relatively easily and is typically lustrous (shiny), malleable, and a good conductor of heat and electricity. Most elements are metals. All from the beginning (left) of the periodic table to the stairs are metals.
    - **Nonmetals** – the relatively small number of elements that appear in the upper-right corner of the period table (to the right of the stairs)

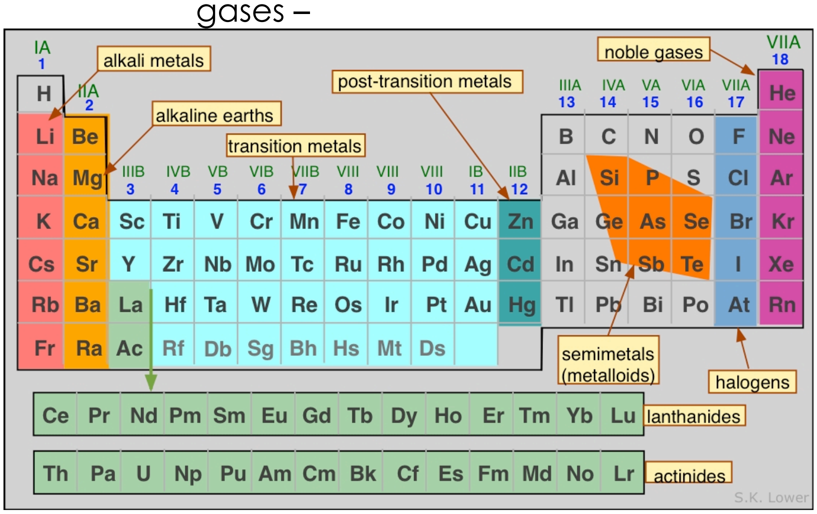
are called nonmetals. Nonmetals generally lack those properties that characterize metals and show much more variation in their properties than metals do. Whereas all metals are solids at normal temperatures, many nonmetals are gaseous, bromine is liquid, and several others are solids.

  - **Metalloids (semimetals)** – the elements that lie close to the stairs line that often show a mixture of metallic and nonmetallic properties. They include silicon, germanium, arsenic, antimony, and tellurium.

- **<u>Properties of Metals and Nonmetals</u>**

Metals	**Nonmetals**
1. Good conductors of heat and electricity (Silver and Copper are the best conductors)	1. Poor conductors of heat and electricity
2. Malleable – can be hammered into sheets	2. Brittle
3. Ductile – can be pulled into wires	
4. Luster (shiny)	3. Dull
5. Solid at room temperature (*except Hg – Mercury)*	4. State varies – Bromine (Br) is a liquid, other are solids and gases
6. Tend to lose/give up electrons	5. Tend to gain electrons

- Know location on table:
  - Alkali metals, alkaline earth metals, transition metals, halogens, noble gases –



- **Symbol** – a letter or letters representing an element of the periodic table. Examples: O, Mn.

5. Separation methods for mixtures

Separation of Mixtures Chart

#	Type of Mixture	Method of Separation	Example	(Type of) Matter
1	Insoluble solid in liquid	Filtering	sand and water	Heterogeneous Mixture
		Centrifuge	blood	Heterogeneous Mixture
2	Liquids immiscible	- Separatory funnel - Decanting	Oil + H_20 (Oil and water)	Heterogeneous Mixture
3	Soluble solid in liquid	- Simple distillation - Evaporation	salt in H_20	Homogeneous Mixture
		Chromatography	ink - marker	Homogeneous mixture
4	2 miscible liquids	Fractional distillation	- alcohol + H_20 - fossil fuels	Homogeneous mixture

***Miscible (mixable)– referring to the ability of two liquids to mix with one another.**

6. Physical/chemical properties; changes; extrinsic and intrinsic properties; energy✓

Properties

Physical properties – properties that can usually be observed with our sense. They include everything about a substance that can be noted when no change is occurring in the type of structure that makes up its smallest component. Some common examples are physical state, color, odor, solubility in water, density, melting point, and hardness.

Chemical properties – properties that can be observed in regard to whether or not a substance changes chemically, often as a result of reacting with other substances. Some common examples are: iron rusts in moist air, nitrogen does not burn, gold does not rust, sodium reacts with water, silver does not react with water, and water can be decomposed by an electric current.

Changes

The changes matter undergoes are classified as either physical or chemical.

Physical change – in general **alters some aspect of the physical properties of matter, but the composition remains constant.** The most often altered properties are **form** and **state**.

- Some examples of physical changes are breaking glass, cutting wood, melting ice, and magnetizing a piece of metal. In some cases, the process that caused the change can be easily reversed and the substance regains its original form. **Water changing its state is a good example of physical changes.** In the solid state, ice, water has a definite volume but takes the shape of the container. When water is heated above its boiling point, it changes to steam. Steam, a gas, has neither a definite size, because its fills the containing space nor shape, because it takes the shape of the container.

Chemical changes – changes in the composition and structure of a substance. They are always accompanied by energy changes. If the energy released in the formation of a new structure exceeds the chemical energy in the original substances, energy will be given off, usually in the form of heat or light or both. This is called an **exothermic reaction**. If however the new structure needs to absorb more energy than is available from the reactants, the result is an **endothermic reaction.**

- Where is energy in the chemical reaction?
  - Exothermic – energy released
  - Endothermic – energy absorbed

- **Examples of chemical changes -** Combustion, corrosion, rusting, acid and base reactions

- **List of "Evidence of chemical reactions"**
  - Color – forming new color
  - Energy change
    - Heat given off (exothermic reaction)
    - Heat absorbed (endothermic reaction)
    - Light Energy - Flash of light
    - Sound Energy
  - Odor – New odor. Gas is evolved (given off), bubbling/fizzing.
  - Formation of a Precipitate – solid formed from combining two solutions
- **Reactants → Products**: Chemical changes are called **reactions**. Silver tarnishes by reacting with substances in the air.
- Law of Conservation of Mass and Law of Conservation of Energy
  - **Law of Conservation of Mass** – the observation that the total mass of materials is not affected by a chemical change in those materials.
  - **Law of Conservation of Energy** – Energy can be neither created nor destroyed, so that the energy of the universe is constant.

Extrinsic and intrinsic properties

- **Also known as extensive/intensive**
  - **Extensive (extrinsic) properties – depend upon quantity of substance**
    - Volume of substance
    - Mass of substance
    - Energy of substance (amount)
  - **Intensive (intrinsic) properties – do not depend upon quantity of substance**
    - Temperature
    - Density ex. A density = 2.70 g/cm^3
    - Melting/boiling point

Energy $q = s\ m\ \Delta T$

- Energy –the ability to do work
- Formula for calculating energy (heat required)
  - Q = energy (heat) required
    - Positive #= heat added/absorbed (endothermic)
    - Negative # = heat released (exothermic)
  - s = specific heat capacity
  - m = mass of the sample in grams

- Δ T = Change in temperature (°C)
  - Final – Initial
- Heat energy, change in temperature (final – initial), heat capacity, specific heat capacity, joules, calories
  - Heat – flow of energy due to a temperature difference
  - Change in temperature (final – initial) – if the equation says the temperature is cooled, then energy is released; if it says heated, then the energy is absorbed
  - Specific heat capacity/ specific heat – the amount of energy required to change the temperature of one gram of a substance by one Celsius degree.
  - Calorie – the amount of energy (heat) required to raise the temperature of one gram of water by one Celsius degree. Joule – an SI unit that can be most conveniently defined in terms of the calorie: **1 calorie (cal) = 4.184 joules (J)**

- Specific heat capacity of metals, specific heat of water – see table below
  - Specific Heat Capacities of Some Common Substances

Substance	**Specific Heat Capacity (J/g °C)**
water (*l*)) (liquid)	4.184
water (*s*) (ice)	2.03
water (*g*) (steam)	2.0
aluminum (*s)*	0.89
iron (*s*)	0.45
mercury (*l*)	0.14
carbon (*s*)	0.71
silver (*s*)	0.24
gold (*s*)	0.13

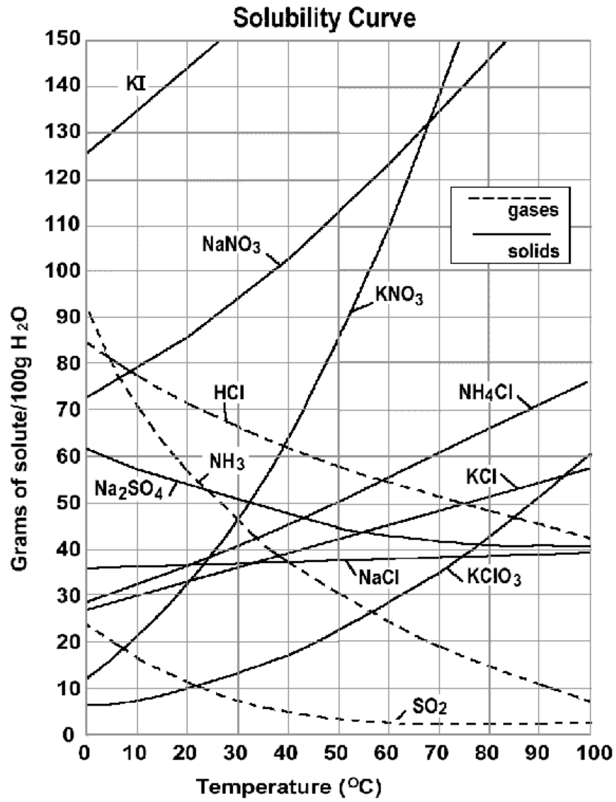
- The symbols *(s), (l), and (g)* indicate the solid, liquid, and gaseous states, respectively.

7. Solubility & solubility curves ✓ - **See solubility chart and p.197**

- **Units of solubility: # grams of solute/100 grams of solvent – See packet with the "Solubility Curve" on the front**
  - solubility = grams of solute/100 grams of solvent
- **Compare solubility of gas and solid solute as temperature increases or decreases)** – see the chart below: Factors that Affect Solubility
- Soluble substance, insoluble substance, saturated solution, unsaturated solutions, supersaturated solutions
  - **Soluble substance** (solid) – a solid that readily dissolved in water
  - **Insoluble substance** (solid) – a solid where such a tiny amount of it dissolved in water that it is undetectable by the human eye
  - **Saturated solution** – solution that has the maximum amount of solute at a given temperature.
  - **Unsaturated solution** – a solution in which more solute can be dissolved than is dissolved already.
  - **Supersaturated solution** - an unstable solution containing more of the solute than it (the solvent) can permanently hold. A rare exception under extremely perfect conditions
- **Solubility –** the ability of solute that can dissolve in a given amount of solvent at a given temperature and pressure.
  - Solution = solute + solvent
    - Solvent - the liquid in which a solute is dissolved to form a solution. **The substance present in the largest amount.**
    - Solute – the minor component in a solution, dissolved in the solvent. Generally a solid (**other substance or substances).**
    - Aqueous Solution – solutions with water as the solvent

Factors that Affect Solubility Chart -

Solute	**Solvent**	**Solubility**
Solids	H_20	• **Temperature –** **increase in temp. = increase in solubility** ○ **sometimes solubility decreases**
Liquids	H_20	• Chemical composition of solute (nature) • **Miscible/immiscible depending on substance**
Gases	H_20	• **Pressure increase = increase in solubility of (dissolved) gas** **Temperature decrease = decrease in solubility**



8. Atomic structure history; discovery of parts of the atom (Dalton, Thompson, Rutherford) – **Topic 3-I SG**

Know

- **Dalton's 5 ideas about atoms**
- **The electron was discovered by J.J Thomson**
- **Rutherford's experiment using alpha particles confirmed that there was mostly empty space between the nucleus and electrons**

<u>Dalton's atomic model (1803)– Billiard Ball Theory – The Beginning of the Modern Atomic Model</u>

- **Proposed the 1st Atomic Theory in 1803**
- It was not until the 18th century that experimental evidence in favor of the atomic hypothesis began to accumulate. Finally, around 1805, John Dalton proposed some basic assumptions about atoms based on what was known through scientific experimentation and observation at the time. These assumptions are very closely related to what scientists presently know about atoms. For this reason, Dalton is often referred to as **the father of modern atomic theory**.

- Some of these basic ideas were: (**some are not true today)**
    1) All matter is made of very small, discrete particles called atoms. (No internal structure, just solid matter)
    2) All atoms of an element are alike in weight, and this weight is different from that of any other kind of atom.
        - **Not true today –** isotopes: have different numbers of neutrons
    3) Atoms cannot be subdivided, created, or destroyed.
        - **Not true today** – can split atoms, they are not indivisible
    4) Atoms of different elements combine in simple whole-number ratios to form chemical compounds
    5) In chemical reactions, atoms are combined, separated or rearranged. (bonds broken and formed)
    - (An atom is **neutral)**

Dalton's Atomic Theory

1. Elements are made of tiny particles called **atoms.**
2. All atoms of a given element are identical.
3. The atoms of a given element are different from those of any other element.
4. Atoms of one element can combine with atoms of other elements to form compounds. A given compound always has the same relative numbers and types of atoms.
5. Atoms are indivisible in chemical processes. That is, atoms are not created or destroyed in chemical reactions. A chemical reaction simply changes the way the atoms are grouped together.

- Dalton used his model to predict how a given pair of elements might combine to form more than one compound. Since he correctly predicted the formation of multiple compounds between two elements, his atomic theory became widely accepted.

- By the second half of the 1800s, many scientists believed that all major discoveries related to the elements had been made. The only thing left for young scientists to do was refine what was already known. This came to a surprising halt when **J.J. Thomson discovered the electron beam in a cathode ray tube in 1897.** Soon afterward, Henry Becquerel announced his work with radioactivity, and Marie Curie and her husband, Pierre, set about trying to isolate the source of radioactivity in her laboratory in France.

C. Composition of The Atom

1. **Matter's electrical connection**

- From around the beginning of the 20th Century, scientists have been gathering evidence about the structure of atoms and fitting the information into a model of the atomic structure.

a. anode/cathode rays; photoelectric effect

<u>JJ Thomson - Anode/Cathode Rays</u>

The discovery of the electron (means beam) as the 1st subatomic particle is credited to **J.J. Thomson** (England, 1897). He used an evacuated tube connected to a spark coil as shown in the figure below. As the voltage across the tube was increased, a beam became visible. This was referred to as a cathode ray. Thomson found that both electrical and magnetic fields deflected the beam. Therefore, he concluded that cathode rays are made up of very small, negatively charged particles, which became known as
electrons.

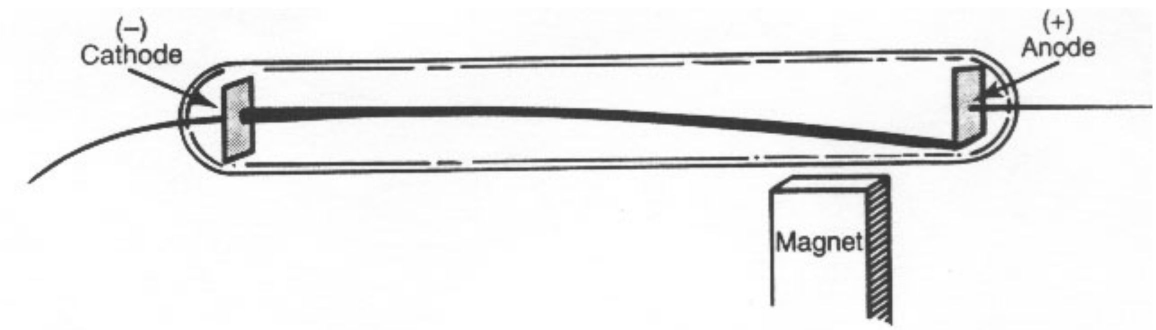
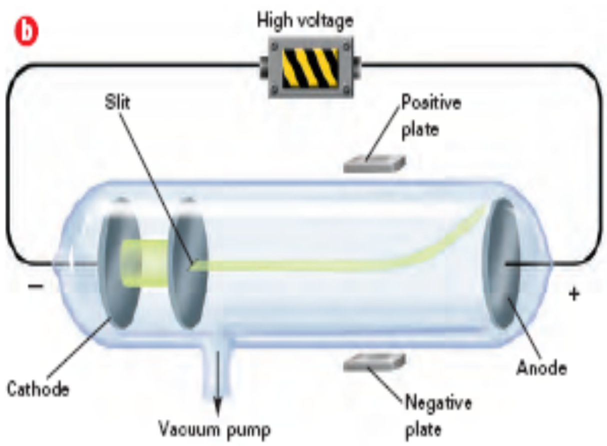
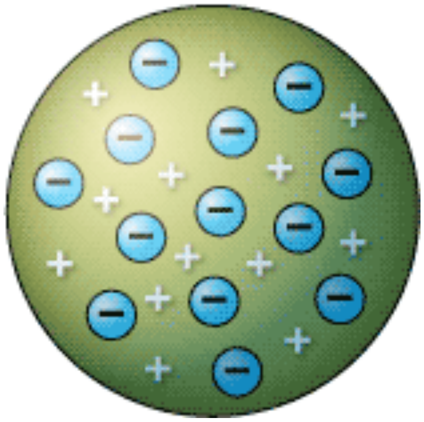


FIGURE 3. *Electron Experiment*



- **He concluded that atoms must also contain positive particles besides negative particles, giving the atom a zero (neutral) overall charge.**
- **Plum Pudding Model (Chocolate Chip Cooke)**
    - **Negative – plums/chips**
    - **Positive – uniform pudding/cookie**



- Further experimentation led Thomson to find the ratio of the electrical charge of the electron to its mass (**charge to mass ratio – is much greater for electrons than protons**). This was a major step toward understanding the nature of the particle.
- "Charge" Unit is a **Coulomb (c)**
    - + or –
        - + - attract
        - ++ **and** -- repel
- electron "e-" is what moves to create electricity

<u>The Nucleus: Rutherford – Gold Foil Experiment</u>

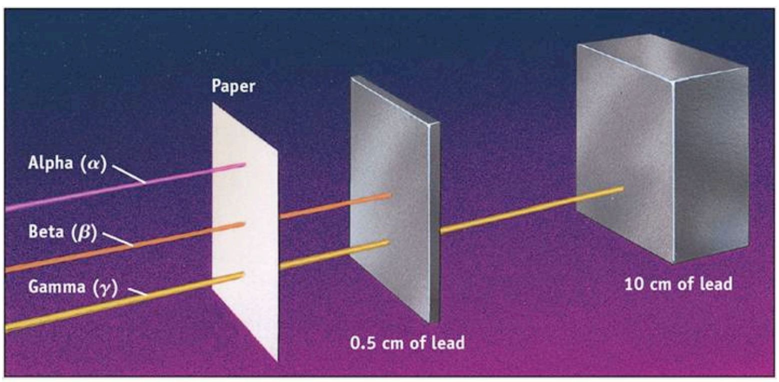
a. discovery of radioactivity; types of radiation **[4.2a]**

- **Origin of radiation = nucleus**
- **Light = radiation**
- **3 types of radioactive emissions**
    - α – alpha particles: charge 2x that of electron, opposite charge
    - β – beta particles: high-speed electron
    - ɣ - gamma rays: high energy "light"

Radioactivity

In the late nineteenth century, scientists discovered that certain elements produce high-energy radiation. For example, in 1896 the French scientist Antoine Henri Becquerel accidentally found that the image of a piece of mineral containing uranium could be produced on a photographic plate in the absence of light. He attributed this phenomenon to a spontaneous emission of radiation by the uranium, which his student, Marie Curie, called **radioactivity.** Studies in the early twentieth century demonstrated three types of radioactive emission: gamma (γ) rays, beta (β) particles, and alpha (α) particles. A γ ray is high-energy "light"; a β particle is a high-speed electron; and an α particle has a 2+ charge—that is, a charge twice that of the electron and with the opposite sign. The mass of an α particle is 7300 times that of

an electron.



b. **Rutherford's experiment**

- **Ernest Rutherford** (England, 1911) performed a gold foil experiment (see below) that had tremendous implications for atomic structure.
- Alpha particles (helium nuclei) passed through the foil with few deflections. However, some deflections (1 per 8,000) were almost directly back toward the source. This was unexpected and suggested an atomic model with mostly empty space between a **nucleus**, in which most of the mass of the atom was located and which was positively charged, and the electrons that defined the volume of the atom. After two years of studying the results, Rutherford finally came up with an explanation. He reasoned that the rebounded alpha particles must have experienced some powerful force within the atom. And he assumed this force must occupy a very small amount of space, because so few alpha particles had been deflected. He concluded that the force must be

a densely packed bundle of matter with a positive charge. He called this positive bundle the nucleus. He further discovered that the volume of a nucleus was very small compared with the total volume of an atom. If the nucleus were the size of a marble, then the atom would be about the size of a football field. The electrons, he suggested, surrounded the positively charged nucleus like planets around the sun, even though he could not explain their motion.

- Further experiments showed that the nucleus was made up of still smaller particles called **protons**. Rutherford realized, however, that protons by themselves could not account for the entire mass of the nucleus. He predicted the existence of a new nuclear particle that would be neutral and would account for the missing mass. In 1932, James Chadwick (England) discovered this particle, the **neutron**. Today the number of subatomic particles indentified and named as discrete units has risen to well over 90.

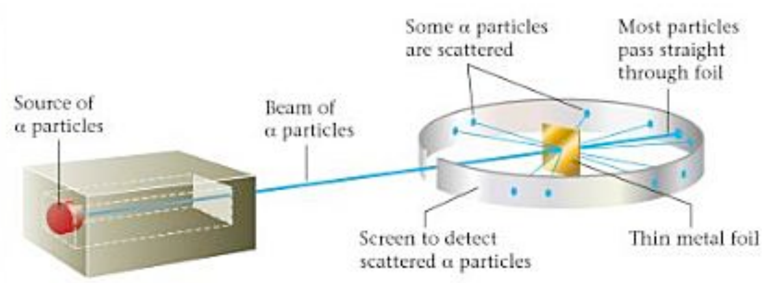


Figure 4.5
Rutherford's experiment on α-particle bombardment of metal foil.

(radium is source)

c. Rutherford's atomic model

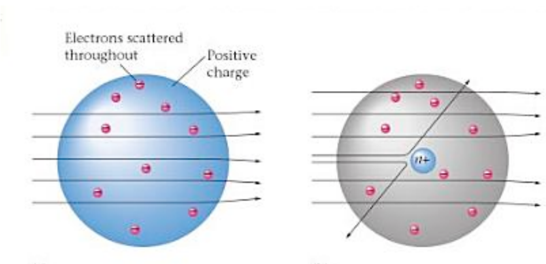
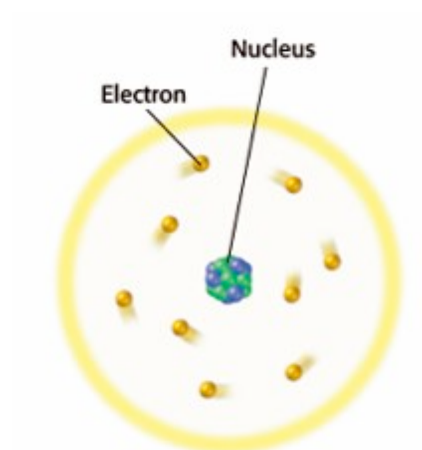


Figure 4.6
(a) The results that the metal foil experiment would have yielded if the plum pudding model had been correct. (b) Actual results.



9. Nucleus & nuclear chem.; **balancing nuclear equations**; transmutations, neutron absorption in nuclear reactors - **Topic III-4**

   - balancing nuclear equations

<u>Honors Chemistry -Outline Nuclear Chemistry Test</u>

Atoms - The Building Blocks of Matter

- **Half-life- time required for half of the atoms of a radioactive nuclide to decay.**
- **5 kinds of emissions (radioactivity)**
  1. **Alpha particles – α (Helium Nuclei)**

$^{4}_{2}He$

Ex.

$$^{238}_{92}U \rightarrow \, ^{234}_{90}Th + \, ^{4}_{2}He$$

- Cannot penetrate paper, skin (not large, so doesn't go very far)
- Very dangerous if ingested

2. **Beta Particles- β or** $^{0}_{-1}e$

Ex. Carbon-14 ($t_{1/2}$ about 5270 years=half life)

$$^{14}_{6}C \rightarrow \, ^{14}_{7}N + \, ^{0}_{-1}e$$

$$^{1}_{0}n \rightarrow \, ^{1}_{1}p + \, ^{0}_{-1}e$$

*What happens in beta emission

- Can penetrate paper, skin
- Can be stopped by thin piece of foil or wood

- artificial transmutation
    - Particle accelerator – accelerates charged particle “bullets” to hit target nuclei. The target nuclei are changed to nuclei of a different element.
- neutron absorption in nuclear reactors?

10. Spectroscopy; light energy-waves & particles ✓ **Topic 3-2**
11. H-atom & quantum theory: Bohr, Heisenberg’s Uncertainty Principle, Pauli’s Exclusion Principle, de Broglie, Aufbau Principle – **Topic III-3**
12. Modern atom; atomic orbitals; electron configurations, energy levels, sublevels, valence electrons; periodic table (families and periods) **Topic III-3 (P.7 of SG)**
13. Atomic masses (amu units) & average atomic masses ✓

p.206-207 text – nuclear chemistry

- Since the actual masses of subatomic particles and atoms themselves are very small numbers when expressed in grams, scientists use **atomic mass units (amu)** instead. An atomic mass unit is defined as ½ the mass of a carbon atom. Thus, the mass of any atom is expressed relative to the mass of one atom of carbon-12.
- The sum of the # of protons and the # of neutrons in the nucleus is called the mass number.
- Average atomic mass – the weighed average of the atomic masses of the naturally occurring isotopes of an element.

14. Law of Definite Composition; % composition ✓

- Definite Composition – a compound is composed of two or more elements chemically combined in a definite ratio by weight.
- Percentage Composition – what percent of the total weight of a compound is made up of a particular element.
    - This simple formula for this is:

$$\frac{\text{Total amu of the element in the compound}}{\text{Total formula amu}} \times 100\% = \text{Percentage composition of that element}$$

To find the percent composition of calcium in calcium hydroxide, we set the formula up as follows:

$$\frac{\text{Ca} = 40 \text{ amu}}{\text{Formula mass} = 74 \text{ amu}} \times 100\% = 54\% \text{ Calcium}$$

To find the percent composition of oxygen in calcium hydroxide:

$$\frac{\text{O} = 32 \text{ amu}}{\text{Formula mass} = 74 \text{ amu}} \times 100\% = 43\% \text{ Calcium}$$

To find the percent composition of hydrogen in calcium hydroxide:

$$\frac{\text{H} = 2.0 \text{ amu}}{\text{Formula mass} = 74 \text{ amu}} \times 100\% = 2.7\% \text{ Calcium}$$

15. Molecular masses ✓ - the mass in grams of one mol of a substance

- By using the atomic masses assigned to the elements, we can find the **formula mass** of a compound. If we are sure that the formula represents the actual makeup of one molecule of the substance, the **term molecular mass** may be used as well. The simplest ratio formula is called **the empirical formula**, and the actual formula is **true or molecular formula.**
- The formula mass is determined by multiplying the atomic mass units (as a whole #) by the subscript for that element and then adding these values for all of the elements in the formula. **See p.127 for example**

16. **The mole ✓; Avogadro's Number ✓**
17. Avogadro's Hypothesis; molar volume; density of a gas at STP ✓; STP
18. Law of Multiple Proportions - states that when two elements combine with each other to form more than one compound, the weights of one element that combine with a fixed weight of the other are in a ratio of small whole numbers. By Dalton.

19. Oxidation numbers; writing formulas -**Topic 4, Part 1**

Writing formulas: Oxidation numbers & polyatomic ions
<u>Formulas with Oxidation Numbers</u> (pgs. 123, examples 124)

- To keep track of the transfer of electrons in all formulas, chemists have devised a system of electron bookkeeping called **oxidation states** (or **oxidation numbers**). In this method, an oxidation state is assigned to each member of a formula or polyatomic ion. It is designed by a small, whole-number superscript ***preceded*** by a plus or minus sign. This is not to be confused with the ionic charges we have been using thus far that are used to the right of the ionic charge. These charges are directly related to the bonding that occurs in compounds. (Oxidations states are also used to track electron transfers.)
- The atom of an electron w/ a stronger pull for electrons is assigned the more negative charge.
- **Rules for Assigning an Oxidation State –** like little sheet
  1. Below are the basic rules for assigning an oxidation state to each element. By applying simple rules, oxidation states can be assigned to most elements or compounds. To apply these rules, remember that the **sum of the oxidation states must be zero for an electrically neutral compound**. For an ion, **the sum of the oxidation states must equal the charge of the ion.**

  The oxidation state of….

  1. An atom in an element is zero. Examples: O for Na(s), O_2(g), and H(l)
  2. A monoatomic ion is the same as its charge. Examples: Na^{+1}, Cl^{-1}
  3. Oxygen is usually -2 in its compounds. Example: H_2O where 2H(+1) + 1O(-2) = 0. (Exceptions occur such as peroxide (O_2^{2-}) when the oxidation state/# is -1.
  4. Hydrogen is usually +1. Examples: H_2O, HCl, NH_3 (The exception is in binary metal hydride compounds like NaH or CaH_2.)
  5. In binary compounds, the more electronegative element is assigned an oxidation number equal to its normal anion charge. Examples: PF_3 : F=-1 and P_2S_3 ; S=-2

<u>Formulas with Polyatomic Ions (page 121 for examples)</u>

- When writing formulas using polyatomic ions, the rules do not change. Simply treat the polyatomic ion as if it were a single anion. If the cation is from category I, follow the rule for category I. If the cation is from category II, follow the rules for category II. The crisscross method does not change either.

20. Nomenclature (naming compounds) (including acids & bases) – **Topic 4, Part 1**

21. Molarity, ✓ molality; and solution concentration

Topic 4-2

22. Stoichiometry of a chemical reaction✓ **- end of topic 4, just learned; see example on pg. 26**

✓ Topic includes possible calculations

Bring to the exam: **Several sharpened pencils and an eraser.**

A nonprogrammable calculator

Something to read if you finish early