

**Chemistry 1201 Review/Preview  
Chapter One Review Guide**

Dr. Sandra McGuire

Spring 2007

**Introduction: Matter and Measurement**

**I. Basic Definitions**

Matter, substance, mixture, physical property, chemical property, intensive property, extensive property, physical change, chemical change, element, compound, mixture, atom, molecule, ion, isotope, etc.

**II. Elements and the Periodic Table** - Know the names and symbols of the first twenty elements, and any additional ones commonly encountered in your class. Know the diatomic elements (hydrogen, nitrogen, oxygen, fluorine, chlorine, bromine, iodine) group names for Group 1A, 2A, 7A, 8A, and the location of the transition metals and semimetals.

**III. Units of Measure and Conversions**

A. SI Units: meters, kilograms, seconds, Kelvins. Know all prefixes from  $10^{15}$  to  $10^{-15}$ .

B. Conversions: English to metric, metric to English, compound units such as miles/hour to feet/second and grams/cm<sup>3</sup> to kg/m<sup>3</sup>.

**IV. Scientific Notation**

$N \times 10^n$  where N is a number between 1 & 10 (for standard notation).

**V. Significant Figures and Calculations**

A. Rules for zeroes (see pages 22 - 23 in text)- leading zeroes are never significant; captive zeroes are always significant; trailing zeroes may or may not be significant.

B. Multiplying and Dividing - the answer has no more significant digits than the number with the fewest significant digits (the least precise figure). Round off after calculations have been performed. (See page 24 for rounding rules.)

C. Adding and Subtracting - The answer has no more decimal places than the addend or subtrahend with the fewest number of decimal places. Note that the number of significant figures is irrelevant in adding and subtracting, it is the decimal places that count

D. Accuracy vs. Precision - Accuracy refers to the agreement of a particular value with the true value. Precision refers to the agreement among several measurements of the same quantity. Systematic error will lead to low accuracy even though precision may be high. Random errors will lead to measurements that are neither accurate nor precise.

**VI. Solving Problems by the Dimensional Analysis**

Include all units on quantities when performing calculations, and cancel appropriate units. (Although this is a good method for solving problems, it is important to THINK through each step rather than simply relying on units to cancel!)

**VII. Temperature** - Know the conversions between the Celsius, Fahrenheit, and Kelvin temperature scales. See Section 1.4 of the text.

**VIII. Density** - a measure of the compactness of matter. Density = Mass/Volume. Some common density units are g/ml (solids and liquids) and g/L (gases).

**NOTE:** *This review sheet is to be used in conjunction with all Chem 1201 lecture notes, handouts, problem sets, the textbook, and other material from Chem 1201.*

**Chemistry 1201 Review/Preview  
Chapter Two Review Guide**

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**Atoms, Molecules, and Ions**

**I. Fundamental Laws of Chemistry**

(Know how to state these and illustrate their operation)

- A. Law of Conservation of Mass
- B. Law of Constant Composition
- C. Law of Multiple Proportions

**II. Experiments Relative to Atomic Structure and Atomic Particles**

(Know how to explain each experiment and what each demonstrated)

- A. The JJ Thomson cathode ray tube experiment
- B. The Milliken Oil Drop Experiment
- C. Rutherford's Experiment

**III. Atomic Structure**

- A. Atomic particles -- protons, neutrons, electrons (mass and charge relationships)  
See Table 2.1 in text.

B. Isotopes:

A	A = mass number ( protons + neutrons)
X	X is the elemental symbol
Z	Z = atomic number (# of protons)

$$\text{Number of neutrons} = A - Z$$

NOTE: All atoms of an element have the same number of protons; only the number of neutrons can change (for different isotopes) or the number of electrons can change (in the case of an ion of the element).

- C. Calculating average atomic weight from weights of individual isotopes

$$\text{Av. At. Mass} = (\text{wt. of isotope 1}) \times \frac{(\% \text{ of isotope 1})}{100} + (\text{wt. of isotope 2}) \times \frac{(\% \text{ of isotope 2})}{100} + \dots$$

**IV. The Periodic Table**

The modern periodic table is arranged according to increasing atomic number, and consists of **vertical groups** (or families) and **horizontal periods**.

The elements in the table are separated into three categories: metals, nonmetals, and metalloids.

Some of the groups have special names. (See Table 2.3)

Group IA	Alkali Metals
Group IIA	Alkaline Earth Metals
Group VIA	Chalcogens
Group VIIA	Halogens
Group VIIIA	Noble Gases

The metals are separated from the nonmetals by a stair step division line. Those elements touching the line are the metalloids, except for aluminum, which is a metal.

**V. Molecules and Molecular Compounds**

Compounds are groupings of elements held together by covalent or ionic bonds.

Covalent compounds (and some polyatomic elements) are composed of molecules; ionic compounds are made up of ions.

*Molecular formulas* indicate the actual numbers and types of atoms in a molecule.

*Empirical formulas* indicate the ratio of the numbers of atoms in a molecule.

e.g. The empirical formula of  $C_2H_{12}$  is  $CH_6$ .

## **VI. Ions and Ionic Compounds**

Ionic compounds are generally formed between metals and non-metals (i.e. the elements from groups I or II with elements from groups VIA and VII A). Some of the ions in ionic compounds are polyatomic ions. You should memorize the ones in Section 2.8 of the text. (Know their formulas, charges, and names.)

Cations are positively charged ions; anions are negatively charged ions.

Cations have fewer electrons than protons; anions have more electrons than protons.

## **VII. Naming Inorganic Compounds -**

Chemical compounds are named systematically depending on the type of compound. See section 2.8 in the text for a discussion of the rules presented below.

### **A. Binary Ionic Compounds in Which the Metal Forms Only One Type of Cation**

These compounds consist of two elements (a metal and a non-metal). The metal forms only one type of cation. These are generally metals from Group IA or IIA (plus silver and aluminum). The cation is named first, followed by the anion whose ending has been changed to -ide. e.g.  $NaCl$  - sodium chloride or  $CaCl_2$  - calcium chloride.

### **B. Binary Ionic Compounds in Which the Metal Can Form More Than One Type of Cation**

In these compounds the charge on the cation is specified by using a Roman numeral. (See Table 2.4 of the text.) Other than that, the naming is similar to that of the previously discussed binary ionic compounds. The most common elements that form more than one type of cation are the transition metals. (See pages 61 and 62 in the text.)

### **C. Ionic Compounds with Polyatomic Ions**

Named according to B or C above, naming the polyatomic ion. You should memorize the names, formulas, and charges of the common polyatomic ions in the text.

### **D. Names and formulas of acids (compounds that contain hydrogen and produce $H^+$ when dissolved in water.**

1. Acids that do not contain oxygen are named with the prefix hydro- and the suffix -ic.  
e.g. hydrochloric acid ( $HCl$ ), hydroiodic acid ( $HI$ ), etc.
2. Acids that do contain oxygen are known as oxyacids and are named according to the number of oxygens in the molecule. See page 66 in the text for a detailed discussion of the rules.

### **E. Names and formulas of binary molecular compounds. Use prefixes to indicate the number of atoms of each type of element in the compound. Note that mono is never used in conjunction with the first element. Learn the prefixes in Table 2.6 in the text.**

## **VIII. Simple Organic Compounds**

Learn the names of the alkanes and alkane derivatives on pages 67-68 of the text.

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**Chemistry 1201 Review/Preview  
Chapter Three Review Guide**

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**Stoichiometry: Calculations with Chemical Formulas and Equations**

**I. Balancing Chemical Equations:**

- A. Memorize the diatomic elements: hydrogen, oxygen, fluorine, chlorine, bromine, iodine
- B. Write the correct chemical formula for all compounds using the charge criss-cross method discussed in Chem 1201 Review/Preview session.
- C. Make the number of each kind of element the same on both sides of the equation by putting in the correct coefficients. Never change the subscripts to balance an equation.

See section 3.1 in the text for more information and examples.

**II. Simple Patterns of Chemical Reactivity**

Three simple types of reactions are frequently encountered in this chapter: combination (or synthesis) reactions, decomposition reactions, and combustion reactions. An example of each type is shown below.

1. *Combination (or synthesis) reactions* are those in which two or more substances combine to form a larger substance. The general representation of a combination reaction is:  $A + B \rightarrow C$
2. *Decomposition reactions* are those in which a compound breaks down into two or more simpler substances. The general representation of a decomposition reaction is  $AB \rightarrow A + B$ .
3. *Combustion reactions* are those that produce a flame. The combustion reactions most commonly seen in this chapter involve a reactant containing C and H (and sometimes O) reacting with oxygen gas,  $O_2$ , to form water and carbon dioxide. (See pages 87-88 of the text for examples.)

**III. Formula Weights**

Formula weights are calculated by adding the atomic weights of all of the atoms in a compound. e.g. The formula weight of  $Ca(NO_2)_3$  is 164.1 amu (atomic mass units). Note that the smallest particles of covalent compounds are called molecules, whereas the smallest units of ionic compounds are called formula units. Therefore, formula weights are often called molecular weights when molecular compounds are involved.

To get the *percentage composition* from formulas, we calculate the percent of each element present, as follows:

$$\% \text{ Element X} = \frac{\text{atomic weight of the element X number of atoms of that element}}{\text{Formula weight of the compound}} \times 100\%$$

**IV. Avogadro's Number and the Mole**

A mole is the amount of substance whose weight in grams is equal to the atomic (or formula) weight of the substance. One mole contains  $6.02 \times 10^{23}$  (Avogadro's number) particles. Interconversions between grams, moles, and particles can be accomplished as shown below. Note that you cannot go directly between grams and particles without first going through moles!

Grams	$\leftrightarrow$	Moles	$\leftrightarrow$	Atoms Molecules Ions
	(Use the formula wt.)		(Use Avogadro's Number)	

See Section 3.4 of the text for a more thorough discussion and examples.

The *molar mass* of a substance is the number of grams in one mole ( $6.02 \times 10^{23}$  particles) of the substance. The units of molar mass are grams per mole (g/mol). Note that the molar mass in grams per mole equals the formula weight in atomic mass units (amu).

## V. Empirical Formula and Empirical Formula Determination

A. Empirical Formula - The simplest formula which shows the ratio of moles of each element in the compound. Empirical formulas can be calculated from percent composition or from combustion data.

To calculate the empirical formula from the % composition:

1. Assume 100 grams of compound (%'s then become grams)
2. Calculate the moles of each element. (moles = grams/atomic mass)
3. Divide by the smallest number to obtain the subscripts in the empirical formula.

(See Section 3.5 of the text for examples and practice problems.)

To calculate the empirical formula from combustion data:

1. Calculate the mass of C in the compound by converting the grams of  $\text{CO}_2$  to moles of  $\text{CO}_2$ . When we get the moles of  $\text{CO}_2$  we have the moles of C because the moles of C equals the moles of  $\text{CO}_2$ . We get the grams of C by multiplying the moles of C by 12.0 (the atomic mass of C.)
2. Calculate the mass of H by converting the grams of  $\text{H}_2\text{O}$  to moles of  $\text{H}_2\text{O}$ . When we get the moles of  $\text{H}_2\text{O}$  we get the moles of H by multiplying the moles of  $\text{H}_2\text{O}$  by two (because each mole of  $\text{H}_2\text{O}$  contains two moles of H). We get the grams of H by multiplying the moles of H by 1.01 (the atomic mass of H.)
3. Calculate the mass of O by finding the sum of the grams of C and H, and subtracting this number from the total grams of compound (The total grams will be given in the problem.) Do Sample exercise 3.15 in the text (without looking at how they did it!)

## VI. Quantitative Information from Balanced Equations

The coefficients in a balanced equation indicate the numbers of moles (or molecules) of each reactant and product involved in the reaction. For example,  $\text{A} + 2\text{B} \rightarrow 4\text{C}$  means that 1 mole(cule) of A reacts with 2 mole(cule)s of B to produce 4 mole(cule)s of C. Note that the number of moles of reactants does not have to equal the number of moles of products. (The total number of grams must be equal on each side, but not the number of moles.)

The numerical coefficients in the balanced equation are used to determine the relationship between moles (and grams) of substances involved in the reaction.

Use the diagram in Figure 3.13 of the text to solve stoichiometry problems by the mole method.

The sequence of steps is as follows:

Grams of given  $\xrightarrow{\text{A}}$  moles of given  $\xrightarrow{\text{B}}$  moles of unknown  $\xrightarrow{\text{C}}$  grams of unknown

For step A, we use molar mass of the given substance, for step B we use the ratio of the coefficients in the balanced equation, and for step C, we use the molar mass of the unknown substance. Note that you should NEVER use the molar mass of one substance with the grams of another substance!!!)

## VII. Limiting Reactants

The limiting reactant in a problem is the reactant that runs out before the other is used up. For example, if we were making tricycles and had 8 frames and 18 wheels, we could make only 6 tricycles before we would run out of wheels. Therefore the wheels would be the limiting reactant and the frames would be the excess reactant. We would have 2 frames left over after we use all of the wheels.

To do a limiting reactant problem in chemistry

Calculate the moles of one reagent (A) necessary to react with the given number of moles of the other reagent (B). If the calculated amount of A is *larger* than the amount of A given in the problem, Reagent A is the limiting reagent (it will run out first). If the calculated amount of A is *smaller* than the given amount, Reagent A is present in excess, and Reagent B is the limiting reagent.

The maximum amount of product possible (the theoretical yield) is stoichiometrically calculated amount of product possible from the limiting reagent.

(See section 3.7 of the text for specific examples.)

The theoretical yield is the maximum amount of product possible from the available reactants. We can calculate the theoretical yield, and we are usually given the actual yield. The definitions and procedure for calculating the % yield are shown below.

1. Theoretical Yield - the amount of product predicted by the balanced equation based on the amount of the limiting reactant.
2. Actual Yield - the amount of product obtained when the reaction is actually carried out in the laboratory.
3. Per Cent Yield:

$$\% \text{ Yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100\%$$

**IMPORTANT:** Read the Strategies in Chemistry (page 108) before you take your next test or quiz). Also, take the Test Preparation workshop on the Center for Academic Success website at [www.cas.lsu.edu](http://www.cas.lsu.edu).

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**Chemistry 1201 Review/Preview  
Chapter Four Review Guide**

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**Aqueous Reactions and Solution Stoichiometry**

**I. General Properties of Aqueous Solutions**

**A. Definitions**

1. Solutions - homogeneous mixtures
2. Solute - the substance that is dissolved (present in the smaller amount).
3. Solvent - the dissolving substance (present in the larger amount)

**B. Electrolytic Properties**

Electrolytes are substances that when dissolved in water, form ions and produce a solution that will conduct electricity. Strong electrolytes ionize completely and produce a strong current; weak electrolytes ionize partially and produce a weak current.

**1. Strong Electrolytes**

Strong electrolytes, which completely break into ions in water, are the strong acids, the strong bases, and soluble salts. The strong acids that you should know are HCl, HBr, HI, HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>, HClO<sub>4</sub>, and HClO<sub>3</sub>.

The strong bases are the hydroxides of the Group IA and IIA metals (except Be). Soluble salts break into ions when placed in water, and the water molecules surround the ions in a process called solvation. (See pages 122 and 123 of the text for a more complete discussion.)

**2. Weak Electrolytes**

Weak electrolytes, which only partially break into ions in water, are the weak acids and weak bases. The weak acids are those that are not strong. Some common weak bases are NH<sub>3</sub> and other compounds that contain nitrogen, hydrogen, and/or carbon, such as CH<sub>3</sub>NH<sub>2</sub>. When weak electrolytes dissolve in water, the dissociation is written as a double arrow because the process is an equilibrium process. (See page 125 of the text.)

**3. Nonelectrolytes** - Substances that do not produce ions when dissolved in water, and therefore do not produce a solution that conducts electricity. Molecular substances that are not acids and bases are non-electrolytes.

**II. Precipitation Reactions**

Precipitation reactions are those reactions that occur when an insoluble solid is produced when two clear solutions are mixed. The insoluble solid is called a precipitate. In order to correctly write precipitation reactions, you must learn the solubility guidelines discussed on page 127 and shown in Table 4.1 of the textbook!

**A. Exchange (Metathesis) Reactions** are reactions in which the positive and negative ions in two compounds exchange partners, as in the general reaction:



See page 128 of the text and Chem 1201 Review/Preview Session notes for the procedure for correctly writing and balancing metathesis reactions.

There are three types of equations to represent metathesis reactions that occur in solution: molecular equations, complete ionic equations (also called ionic equations), and net ionic equations. The molecular equation shows all species that take part in the reaction, written as if they were molecules. The complete ionic equation shows all strong electrolytes written in ionic form. The net ionic

equation shows only those species that actually change during the course of the reaction. All spectator ions are omitted. Spectator ions are those ions that appear on both sides of the equation. See pages 128 – 130 of the text, and do all Sample Exercises (without looking at the solution!)

### III. Acid-Base Reactions

Remember to memorize the common strong acids and the strong bases! (Table 4.2)  
See Section IB of this review guide for a discussion of how to identify strong electrolytes, weak electrolytes, and non electrolytes.

#### A. Neutralization Reactions and Salts

The general reaction for an acid-base reactions is shown below:



See Sample exercise 4.7 to practice writing and balancing these types of equations.

#### B. Acid-Base Reactions with Gas Formation

When a gas is produced in an acid-base reaction, the gas is always written as a molecule. See page 136 for more information.

### IV. Oxidation-Reduction Reactions

When electrons are transferred between the reactants, the element that loses electrons is said to be oxidized and the element that gains electrons is said to be reduced. Remember, LEO the lion goes GER (**L**oss of **E**lectrons is **O**xidation; **G**ain of **E**lectrons is **R**eduction.)

Oxidation numbers are compared to tell is a reaction is a redox reaction.

#### 1. Assigning Oxidation Numbers

Oxidation numbers represent the apparent charge on a atom when all of the electrons are assigned to the more electronegative atom in the bond. *You must memorize the rules on page 139 of the textbook if you want to correctly assign oxidation numbers!* The rules are not difficult to learn because they are very logical and straightforward.

#### 2. Oxidation of metals by acids and salts

When metals react with acids and salts, a displacement reaction occurs.



The metal loses electrons to X, and is therefore oxidized

3. The Activity Series of the Elements arranges elements in order of their tendency to lose electrons (i.e. to reduce other elements). See Table 4.5 of the text and pages 142 – 144 for a detailed discussion.

### V. Concentrations of Solutions

A. Molarity (M) - moles of solute per liter of solution.

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

Note that this equation can be solved for moles of solute or liters of solution.

When expressing the concentrations of electrolytes, we must remember that they dissociate. Therefore 1 mole of  $\text{Ca}(\text{NO}_3)_2$  will produce 3 moles of ions in solution. (See sample exercise 4.2)

## B. Dilution

When a more concentrated solution is used to make a less concentrated solution, the following relationship is used in the calculations:

$$M_{\text{conc}} \times V_{\text{conc}} = M_{\text{dil}} \times V_{\text{dil}}$$

## VI. Solution Stoichiometry and Chemical Analysis

The stoichiometric calculations involving reactions in solution must be performed based on volumes of solutions rather than on grams of reactants. We must calculate moles from the molarity and volume of solution, using the relationship

$$\text{Moles} = \text{Molarity (moles/L)} \times \text{Volume (L)}$$

After we have calculated the moles of reacting substance, the procedure is the same as we've used before to determine the moles of product.

See Section 4.6 and figure 4.19 for a detailed description and examples.

Titration calculations are often used to determine the concentration of a particular solute in a solution. Titration calculations involve using the balanced equation and stoichiometric relationships to determine the concentration of the solute in question. See pages 152 - 156 in the text for examples of these types of problems.

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**Chemistry 1201 Review/Preview**  
**Chapter Six Review Guide**

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Spring 2007

**Electronic Structure of Atoms**

**I. The Wave Nature of Light**

A. The electromagnetic radiation spectrum illustrates the range and regions of electromagnetic radiation. The spectrum ranges from the high-energy gamma region to the low energy radio wave region and beyond. The visible region of the spectrum ranges from violet (400 nm) to red (750 nm).

B. The relationship between the frequency of electromagnetic radiation, the wavelength of the radiation, and the speed of light is given by the following relationship:  $v\lambda = c$

where  $c$  is the speed of light ( $3 \times 10^8$  m/sec);  $\lambda$  is the wavelength of the radiation; and  $v$  is the frequency.

(Do sample exercises 6.1 and 6.2 (without looking at the solution!) for practice using this equation.)

**II. Quantized Energy and Photons**

Max Planck theorized that energy can be emitted or absorbed by atoms only in discrete chunks, called quanta. One quantum is the smallest amount of energy that can be transferred. The relationship between the energy of a single quantum and the frequency is given by the following relationship:

$$E = hv \quad \text{where } h = \text{Planck's constant, } 6.626 \times 10^{-34} \text{ J}\cdot\text{s}$$

(See Exercises on page 223 for practice.)

**III. Line Spectra and the Bohr Model**

A. Emission spectra (or line spectra) - produced when an excited atom emits energy of specific wavelengths due to the electrons falling from higher to lower energy levels. (See Figure 6.12.)

B. The Rydberg equation - provides the relationship between the wavelength or frequency of a given line in the spectrum for hydrogen and the energy levels involved in the transition as

$$1/\lambda = R_H [1/n_1^2 - 1/n_2^2] \quad \text{or} \quad \nu = R_{HC} [1/n_1^2 - 1/n_2^2]$$

where  $R_H$  is the Rydberg constant ( $1.096776 \times 10^7 \text{ m}^{-1}$ )  
and  $n_1$  and  $n_2$  are integers with  $n_2 > n_1$

C. The energy states of the hydrogen atom and the changes in energy when the electron changes states can be calculated using the equations shown on pages 226 and 227.

The ground state of the hydrogen atom occurs at  $n = 1$ . Energy states of  $n = 2$  or higher are referred to as excited states.

See Section 6.3 of the text for an in depth discussion and practice exercises.

**IV. The Wave Behavior of Matter**

A. The de Broglie equation shows the relationship between the mass, wavelength, and velocity of any moving particle:  $\lambda = h/mv$

See section 6.4 of the text for more information and examples.

B. Heisenberg Uncertainty Principle - It is not possible to know precisely both the position and the speed of an electron at a given point in time. The more precisely one is known, the less precisely the other is known.

**V. Quantum Mechanics and Atomic Orbitals**

The wave function of an electron is represented by  $\Psi$ , while the probability of finding the electron in a given space is given by  $\Psi^2$ . Quantum mechanics uses a

set of four quantum numbers to explain the distribution of electrons in atoms. The four quantum numbers are designated by  $n$ ,  $l$ ,  $m_l$ , and  $m_s$ .

A. The principal quantum number,  $n$ , represents the principal energy level of the electron.  $n$  can have values of 1,2,3,... No elements currently existing have a principal quantum number greater than 7.

B. The azimuthal quantum number (or angular momentum quantum number),  $l$ , represents the sublevel of the electron and defines the three dimensional shape of the orbital.

$l$  can have values of 0, 1, 2, 3, ...  $n-1$ .  $l = 0$  is the s sublevel,  $l = 1$  is the p sublevel,  $l = 2$  is the d sublevel,  $l = 3$  is the f sublevel,  $l = 4$  is the g sublevel, etc. No currently existing elements have electrons in an  $l = 4$  or higher sublevel.

C. The magnetic quantum number,  $m_l$ , represents the orbital in which the electron resides. The values of  $m_l$  are  $-l$  to  $+l$ . Thus, when  $l = 2$ , the values of  $m_l$  are -2, -1, 0, +1, and +2. Each orbital can hold a maximum of 2 electrons.

D. The spin quantum number,  $m_s$ , indicates the spin of the electron. The values of  $m_s$  are  $+1/2$  and  $-1/2$ .

## VI. Representations of Orbitals

The shapes of the atomic orbitals are given on pages 235-237 of the text. There is one spherically shaped s orbital in each subshell; there are three dumbbell shaped p orbitals in each subshell (beginning with the second principal energy level); beginning with the third energy level there are five d orbitals in each subshell (four of them are double dumbbells lying in a plane and the fifth is a single dumbbell in a doughnut); and beginning with the fourth energy level there are seven f orbitals.

## VII. Many Electron Atoms

A. Orbitals and their energies

The energy increases as the  $n$  value increases. Within a given energy level ( $n$  value) the energy increases with increasing  $l$  values. (See page 238.)

B. The Pauli Exclusion Principle states that no two electrons in the same atom can have the same four quantum numbers.

## VIII. Electron Configurations

The way in which electrons are distributed in the orbitals of an atom is called the electron configuration.

A. The Building-Up (Aufbau) Principle

The order of electron filling is given by the Aufbau Principle, as illustrated in Section 6.8 of the text. Page 249 presents a discussion of electron configurations of anomalous elements. Read these sections carefully and know the anomalies!

B. Hund's Rule: Subshells containing more than one orbital will half fill each orbital with electrons with paired spins before the orbitals become double occupied with electrons of opposite spins.

C. The Aufbau Principle and the Periodic Table

The members of the families in the main group elements in the periodic table have the same number of valence electrons, and thus the same valence electron configuration. (See Figure 6.30.) The *valence electrons* are the electrons in the principal energy level with the largest value of  $n$ . The valence electrons dictate the behavior of the elements. Vertical groups, or families, in the periodic table have the same number of valence electrons.

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**Chemistry 1201 Review/Preview  
Chapter Seven Review Guide**

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**Periodic Properties of the Elements**

**I. Development of the Periodic Table**

A. Mendeleev's Contribution

When the properties of the elements were examined, Mendeleev discovered that their chemical properties repeated in a regular way when they were arranged according to increasing atomic weights. The repeating behavior of the chemical properties of the elements is called chemical periodicity or periodic behavior.

B. The Modern Periodic Table

After more elements were discovered and their properties studied, it was clear that the correct arrangement of the elements is in order of increasing *atomic number*, not atomic weight. The modern periodic table is arranged according to increasing atomic number, and consists of vertical groups (or families) and horizontal periods.

(See pages 262-263 in the text.)

**II. Effective Nuclear Charge**

A. Effective Nuclear Charge - In multielectron atoms, the electron-to-electron repulsions reduce the strength of the nuclear charge felt by the electrons. The effective nuclear charge,  $Z_{\text{eff}}$ , is the net nuclear charge actually felt by an electron, and is calculated as follows:

$$Z_{\text{eff}} = Z (\text{the number of protons}) - S (\text{the electron shielding constant})$$

B. Periodic Trends in Effective Nuclear Charge

The effective nuclear charge,  $Z_{\text{eff}}$ , increases as we move from left to right across the periodic table, but decrease as we move from top to bottom. (See pages 264-265.)

**III. Sizes of Atoms and Ions**

A. Atomic Radii generally decrease from left to right and increase from top to bottom. (Do Sample Exercise 7.2 without looking at the solution.)

B. Ionic Radii depend on the type and size of the charge. Positive ions (cations) are smaller than their parent atoms, and negative ions (anions) are larger than their parent atoms. If ions have the same charge, the size increases from top to bottom in the periodic table. (See Figure 7.7 in the textbook for a comparison of the sizes of positive and negative ions and their parent atoms.)

C. Isoelectronic Series – a group of ions that contain the same number of valence electrons. E.g.  $\text{Na}^+$ ,  $\text{Mg}^{2+}$ ,  $\text{Al}^{3+}$ ,  $\text{O}^{2-}$ , etc. In any isoelectronic series, the anions will be larger than the neutral atom containing the same number of electrons as the ions in the series, and the cations will be smaller. The greater the negative charge, the larger the anion; the greater the positive charge, the smaller the cation. (See Sample Exercise 7.4 for practice with this.)

*Important note: Pay close attention to whether the problem asks you to arrange in order of increasing or decreasing size!*

**IV. Ionization Energy**

A. Definition of ionization energy

The ionization energy is the energy required to remove an electron from the gaseous atom in the ground state. i.e.  $\text{M} + \text{energy} \rightarrow \text{M}^+$ . The ionization energy increases from left to right when moving across the periodic table,

and decreases from top to bottom. These trends are due to the differences in the effective nuclear charge experienced by the electron being removed.

B. Successive ionization energies

Because more than one electron can be removed from an atom or ion, there can be more than one ionization energy value. The energy required to remove the first electron is  $I_1$ , the energy required to remove the second electron is  $I_2$ , etc. As shown in Table 7.2 the ionization energies always increase as we continue to remove electrons. There is always a huge jump in ionization energy after all of the valence electrons have been removed. E.g.  $I_3$  for Mg is much, much bigger than  $I_2$ . (Do Sample Exercise 7.5 for practice.)

C. Periodic Trends in First Ionization Energies

1. The ionization energies generally increase going across a row. (Why?)
2. The ionization energies generally decrease going down a group. (Why?)
3. The main group elements have a greater range of ionization energies than do the transition metals.

(See pages 272 and 273, and Figure 7.11 for a more complete understanding of these concepts. Be sure to do Practice Exercise 7.6)

D. Electron Configurations of Ions

1. Cation Formation

When electrons are removed from an atom to form a cation, the valence electrons are always removed first. Remember that the valence electrons are those in the highest principal energy level. E.g. The 4s electrons are always removed before the 3d electrons, even though electrons go into the 4s orbitals before they populate the 3d orbitals.

2. Anion Formation

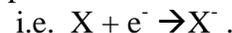
When electrons are added to an atom to form an anion, they always go into the available orbital in the principal energy level with the lowest value of  $n$ .

(Do Sample Exercise 7.7 for practice.)

V. **Electron Affinities**

A. Definition of Electron Affinity

Electron Affinity is the energy associated with gaining an electron in the gas phase.



B. Electron Affinity Values of the Elements

Almost all electron affinity values are negative because of the attraction between the added electron and the nucleus. However, the noble gases have an electron affinity of 0, and a few other elements have a positive electron affinity because the added electron feels a relatively small effective nuclear charge. The halogens have the highest (most negative) electron affinities. (See section 7.5 of the text and your lecture notes for more information on this.)

VI. **Metals, Nonmetals, and Metalloids**

A. Properties of Metals vs. Nonmetals

See Section 7.7 and Table 7.3 on page 277 for the properties of these substances.

B. Metalloids

Metalloids have properties between those of metals and nonmetals. For example, they may have some characteristics of metals, such as shininess, but not others.

## VII. Group Trends for the Active Metals

### A. The Alkali Metals (Group 1A)

The alkali metals are extremely reactive because they readily lose an electron. They will vigorously react with water, to produce hydrogen gas. Also, they react with hydrogen to produce metal hydrides:  $2\text{Na} + \text{H}_2 \rightarrow 2\text{NaH}$  They also react with oxygen, but the reactions are a bit more complex. See pages 282-283 in the text, and Sample Exercise 7.10 for practice in writing the reactions of the alkali metals.

### B. The Alkaline Earth Metals (Group 2A)

The alkaline earth metals are less reactive than the alkali metals, but are still quite reactive with substances such as the halogens, water, and oxygen. (See page 285 for more detail on this topic.)

## VIII. Group Trends for Selected Nonmetals

### A. Hydrogen

Because hydrogen has one valence electron, which it can gain or lose, it can become a cation ( $\text{H}^+$ ) or an anion ( $\text{H}^-$ ). It is a nonmetal, but acts as metals do when it loses its electron. Hydrogen reacts with nonmetals to lose its electrons, and with active metals to gain an electron. The reaction of hydrogen with oxygen to form water is very exothermic.

### B. Groups 6A and 7A

These groups tend to react with nonmetals to take electrons from them.

### C. Group 8A: The Noble Gases

The noble gases are generally unreactive, and are said to be inert. However, since 1962 Xe and Kr compounds containing fluorine have been made in reactions with fluorine containing compounds.

**NOTE: This review sheet is for Chem 1201 Review/Preview attendees and is to be used in conjunction with lecture notes, handouts, problem sets, the textbook, and any other material from Chemistry 1201.**

**Chemistry 1201 Review/Preview  
Chapter Eight Review Guide**

**Dr. Sandra McGuire**

**Spring 2007**

**Basic Concepts of Chemical Bonding**

**I. Chemical Bonds, Lewis Symbols, and the Octet Rule**

- A. Chemical Bond – the strong attraction between two atoms or ions
  - 1. Ionic Bond – the electrostatic attraction between ions of opposite charges
  - 2. Covalent Bond – The bond formed due to sharing of electrons between atoms
  - 3. Metallic bond – The bond found in metals, due to freely moving electrons in the metal.
- B. Lewis Symbols – show the symbol of an atom, along with the valence electrons. (See Table 8.1 for the Lewis symbols of many of the elements.)
- C. The Octet Rule states that atoms will gain, lose, or share electrons in order to obtain eight electrons in their valence shell.

**II. Ionic Bonding**

- A. Energetics of Ionic Bond Formation  
When forming ionic bonds from neutral atoms, and electrons must be removed from the neutral metal atoms to form cations; electrons must be added to the nonmetal atom to form anions; and the cations and anions must be in close enough proximity for the electrostatic attraction to be felt. Therefore, formation of an ionic bond involves different energies: ionization energy, electron affinity, and lattice energy. The ionization energy is positive (forming the cation requires energy input), the electron affinity is generally negative (forming the anion generally releases energy), and the formation of the lattice releases energy. (See pages 305 – 306 and lecture notes for a more complete discussion.)
- B. Electron Configuration of Ions of the Representative Elements  
The atoms of elements in groups 1A, 2A, and 3A form cations with charges of +1, +2, and +3, respectively. The atoms of elements in groups 5A, 6A, and 7A form anions with charges of -3, -2, and -1, respectively.
- C. The Transition Metals  
The transition metals lose their s electrons first, and then d electrons. Therefore the charges on transition metal ions are generally +1, +2, or +3, depending on how many s and d electrons are lost.

**III. Lewis Structures**

- A. What are they?  
Lewis structures show the arrangement of electrons in a molecule. The electrons shown are the valence electrons of all of the atoms in the molecule.
- B. Multiple Bonds  
When more than one pair of electrons (two electrons) are shared between two atoms, the bond is called a multiple bond. Double bonds involve the sharing of two pairs of electrons (four electrons), and triple bonds involve the sharing of three pairs of electrons (six electrons).

#### IV. Bond Polarity and Electronegativity

- A. *Bond polarity* describes the way the electrons are shared between the atoms. A bond is a nonpolar covalent bond if the electrons are fairly equally shared between the atoms. A bond is a polar covalent bond if the electrons are shared unequally.
- B. *Electronegativity* is the ability of an atom to attract shared electrons. Electronegativity values for the main group elements are shown in Table 8.6 of the textbook, shown on page 312 of the textbook. You should notice three important aspects about the table. You should commit these to memory.
1. The four most electronegative elements (from highest to lowest) are fluorine, oxygen, chlorine, and nitrogen.
  2. The electronegativity generally increases from left to right across a period, and decreases from top to bottom down a group.
  3. The noble gases have no electronegativity values assigned. This is because electronegativity is a measure of an atom's ability to attract **shared** electrons. Since the noble gases generally do not form compounds (they have a stable octet of electrons), they do not share electrons.

It is the **difference** in electronegativity that determines whether two atoms will form an ionic bond, a nonpolar covalent bond, or a polar covalent bond. If there is a large difference in the ability to attract shared electrons, the more electronegative atom will completely take the electrons from the less electronegative element, resulting in a transfer of electrons, thus forming an ionic bond.

If the electronegativities of the atoms are equal, or nearly equal, the electrons will be equally shared between the two atoms, resulting in a nonpolar covalent bond.

If the electronegativities of the atoms are not equal, but not vastly different, the electrons will be unequally shared between the two atoms, resulting in what is called a polar covalent bond.

#### C. Illustrating Bond Polarity

The partial charges in polar covalent bonds are indicated by placing a  $\delta^+$  sign over the partially positive atom, and a  $\delta^-$  sign over the partially negative atom. Alternatively a  $+ \text{---} \rightarrow$  can be used, in which the arrow is placed over the partially negative atom. Both conventions are shown in the HCl molecule below.



(See pages 314 – 316 of the text for more information.)

#### V. Drawing Lewis Structures

##### A. Rules

The general rules for writing Lewis dot structures are written below. (*You should memorize these rules, or those presented in lecture or the textbook, and practice, practice, practice drawing Lewis structures.*)

1. Determine the total number of valence electrons to be shown in the structure. For negative ions add in the negative charge, for positive ions subtract the positive charge.
2. Write the skeletal structure showing what atoms are bonded together.

3. Assign each outer atom an octet (except H), and count the total number of electrons shown.
4. Place additional electrons around the central atom if the number of electrons shown in the structure is less than the number of electrons determined in step one. (Note that the elements beyond the second period of the periodic table can have more than eight electrons surrounding them. Phosphorus often has ten electrons around it and sulfur often has twelve electrons surrounding it.)
5. If the number of electrons placed in the structure is correct, but an atom that should have an octet does not have one, double (or triple) bonds can be placed by converting a lone pair on one atom to a bonding pair between that atom and the one deficient in electrons.

*Do Sample Exercises 8.6, 8.7, 8.8, and 8.9 for practice drawing Lewis structures.*

- B. Number of Bonds and Lone Pairs Often Encountered for Selected Elements  
The usual number of bonds and lone pairs for different atoms (when the formal charge on the atom is 0) is as shown below. Note that these are common, but there are many exceptions!

H:	1 bond	0 lone pairs
O:	2 bonds	2 lone pairs
N:	3 bonds	1 lone pair
C:	4 bonds	0 lone pairs
F:	1 bond	3 lone pairs

Other halogens: 1 bond 3 lone pairs (when the halogen is not a central atom)

- C. Formal Charge

1. Determining formal charge

The formal charge is the charge on an atom when the electrons in a Lewis dot structure are assigned to individual atoms as follows: each atom is assigned all of its non-bonding electrons (electrons in lone pairs) and half of its bonding electrons.

Formal charge = (the # valence electrons in free atom) minus (the # of non-bonding electrons) minus (1/2 the # of bonding electrons)

2. Using the formal charge to determine the most correct Lewis structure  
The most correct Lewis structure is the one in which:
  - a) the atoms have formal charges closest to zero
  - b) the negative formal charges are on the most electronegative elements

## VI. Resonance Structures

When two or more Lewis dot structures can represent the arrangement of electrons in the same molecule or ion, the individual Lewis structures are called resonance structures. The actual Lewis structure is a combination of all of the contributing resonance structures. (See pages 322 – 325 of for examples and practice.)

## VII. Exceptions to the Octet Rule

- A. Odd electron molecules - Although rare, there are molecules that contain an odd number of electrons. Some examples are  $\text{ClO}_2$  and  $\text{NO}$ . In these molecules, the least electronegative element is usually the one that is written with a deficiency of electrons.

- B. The incomplete octet - B and Be can exist with 6 or 4 electrons, respectively. (H never has more than 2 electrons!)
- C. The expanded octet - As stated above, the elements beyond Period 2, such as P and S can have more than 8 electrons.

### VIII. Strengths of Covalent Bonds

The measure of the strength of a bond is the amount of energy required to break the bond. The *bond enthalpy* is the quantitative description of this. Bond enthalpies are always positive because energy is always required to break bonds.

Table 8.4 provides the average bond enthalpies of some commonly encountered bonds. Notice that single bonds generally have lower bond enthalpies than multiple bonds, and that longer bonds generally have lower enthalpies than shorter bonds.

#### A. Bond Enthalpies and the Enthalpies of Reactions

The enthalpy of a reaction can be determined from the bond enthalpies of the reactants and products as shown below:

$$\Delta H_{\text{reaction}} = \Delta (\text{bond enthalpies of bonds broken}) - \Delta (\text{bond enthalpies of bonds formed})$$

(Do Sample Exercise 8.12 for practice with these types of problems.)

#### B. Bond Enthalpy and Bond Length

In general, the stronger a bond the shorter the bond. As shown in Table 8.5, single bonds are generally longer than double bonds, and double bonds are longer than triple bonds.

**NOTE: This review sheet is for Chem 1201 Review/Preview attendees and is to be used in conjunction with lecture notes, handouts, problem sets, the textbook, and any other material from Chemistry 1201.**

**Chemistry 1201 Review/Preview  
Chapter Nine Review Guide**

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**Adj. Prof., Dept. of Chemistry**

**Spring 2007**

**Molecular Geometry and Bonding Theories**

**I. Molecular Shapes**

The three dimensional shape of a molecule is determined from the number of electron domains (atoms and nonbonding pairs of electrons) connected to the central atom. The steps involved in determining the shape of a molecule are written below:

1. Draw the correct electron dot structure.
2. Count the number of electron domains (atoms plus lone pairs) around the central atom.
3. Determine the electron arrangement around the central atom. (This is the shape of the molecule if there are no lone pairs.)
4. If lone pairs are present, describe the shape of the molecule ignoring the position(s) occupied by any lone pairs

See figures 9.1, 9.2, 9.3, and 9.4 for pictures of the possible shapes of molecules.

**II. The VSEPR Model**

The Valence Shell Electron Pair Repulsion Model, or VSEPR (pronounced “vesper”) is based on the concept that the valence electrons in a molecule will repel each other. The shape that a molecule (or ion) assumes is the one that allows the valence electrons to be as far apart as possible.

**A. Determining the Molecular Shape**

To correctly determine molecular shape we must **always begin with a correct Lewis structure** that will tell us how many valence electrons are in the molecule. (If we start with an incorrect Lewis structure, we will most probably assign the molecule an incorrect shape. Always exercise extreme care when drawing Lewis structures to make sure they are correct.) When we have the correct dot structure, we can determine the number of electron domains (the number of electron groups around the central atoms of the molecule). The steps to follow in determining the shape of a molecule are summarized below.

1. Draw the correct Lewis structure for the molecule
2. Determine the electron domain geometry by spacing the electron pairs as far apart as possible, as shown in the table below.

<u>Pairs of electrons</u>	<u>Electron domain geometry</u>
2	linear
3	trigonal planar
4	tetrahedral
5	trigonal bipyramidal
6	octahedral

The molecular shape associated with each number of groups of electrons (electron domains) is shown in Table 9.1 on page 349 of the text.

**B. The Effect of Nonbonding Electrons and Multiple Bonds on Bond Angles**

When we consider molecules that have nonbonding (lone) pairs of electrons, we determine the shape of the molecule by ignoring the position of the lone pairs.

**C. Molecules with Expanded Valence Shells**

Molecules that have an expanded octet can assume the shapes of trigonal bipyramidal and octahedral. When nonbonding pairs are introduced, the possible molecular geometries is increased, as shown in Table 9.3 on page 354 of the text.

The possible electron domain and molecular shapes (ignoring the position of lone pairs) for molecules with octets and extended octet domains are summarized in Tables 9.2 and 9.3 on pages 351 and 354 of the text.

### III. Molecular Shape and Molecular Polarity

#### A. Determination of Molecular Polarity

Two factors must be considered when determining whether a molecule is polar or non-polar:

1. Whether any polar bonds are present in the molecule, due to differences in electronegativity. If there are no polar bonds in the molecule, the molecule **must** be non-polar. If a molecule is non-polar, we say that the dipole moment is zero.
2. Whether the individual bond dipoles cancel each other, based on the shape of the molecule. (We use vectors to determine whether the individual bond dipoles cancel, as described in Section 9.3, and shown in figures 9.11 and 9.12 of the textbook.) If all of the bond dipoles cancel, the molecule will be non-polar.

For example, the  $\text{CO}_2$  molecule has two polar  $\text{C}=\text{O}$  bonds, but because the molecule is linear, the bond dipoles cancel and the molecule is nonpolar. The dipole moment is therefore zero. However, for the  $\text{OCS}$  molecule, the bond dipoles are not equal and do not cancel each other. Therefore,  $\text{OCS}$  has a nonzero dipole moment.

#### B. Molecules That Will Always Be Nonpolar (Have a Zero Dipole Moment)

The following types of molecules will always be nonpolar.

1. Linear molecules with the general formula  $\text{A}_2\text{X}$
2. Trigonal planar molecules with the general formula  $\text{A}_3\text{X}$
3. Tetrahedral molecules with the general formula  $\text{A}_4\text{X}$
4. Trigonal bipyramidal molecules with the general formula  $\text{A}_5\text{X}$ , or  $\text{A}_3\text{B}_2\text{X}$  where all three equatorial atoms are the same and both axial atoms are the same
5. Octahedral molecules with the general formula  $\text{A}_6\text{X}$ , or  $\text{A}_4\text{B}_2\text{X}$  where the four atoms

Do sample exercise 9.4 for more practice with this concept.

#### C. Covalent Bonding and Orbital Overlap

The valence bond theory describes bonds as being formed from the overlap of orbitals. For example, the  $\text{H}_2$  bond is formed from the overlap of the s orbital on one hydrogen with the s orbital on the other hydrogen. Similarly, the  $\text{F}_2$  bond is formed from the overlap of the 2 p orbital on each atom. The bond forms because the electrons in the overlapped orbitals are attracted to both nuclei, and therefore hold the atoms together. When a bond is formed there is an optimum distance between the two nuclei, and there is a resulting decrease in potential energy. (See Figure 9.15.)

#### D. Hybrid Orbitals

In order to explain the shapes of molecules we invoke the concept of hybridization of atomic orbitals, explained briefly below. (This is covered in detail in Section 9.5 of the text.)

**Hybridization** - mixing of atomic orbitals in an atom to form a new set of "hybridized" orbitals. The hybridization of an atom can be determined from the total number of atoms plus lone pairs (or electron domains) connected to it in the Lewis structure of the molecule. The hybridization for each of the electron domain numbers is shown on the following page. Carefully study Table 9.4 on page 366 for a more thorough understanding of this concept.

<u>Atoms + Lone Pairs (Electron Domains)</u>	<u>Hybridization</u>
2	sp
3	sp <sup>2</sup>
4	sp <sup>3</sup>
5	sp <sup>3</sup> d
6	sp <sup>3</sup> d <sup>2</sup>

## VI. Multiple Bonds

### A. Types of bonds

When more than two electrons (one pair) are shared between atoms, a multiple bond (double or triple bond) is formed. There are two different types of bonds formed from orbital overlap, described below.

1. Sigma bonds - formed from end to end overlap of orbitals
2. Pi Bonds - formed from sideways overlap of unhybridized p orbitals (Note: d orbitals can be involved in pi bonding also, but this is usually not covered in Chem 1201.)

### B. Sigma ( $\sigma$ ) and Pi ( $\pi$ ) Bonding as related to single, double, and triple bonds

In general, single bonds are composed of a sigma bond, double bonds are formed from one sigma bond and one pi bond, and triple bonds are formed from one sigma bond and two pi bonds.

### C. Delocalized $\pi$ bonding

When there are three or more atoms that are all in the same plane and are connected by pi bonds, the electrons can be delocalized as shown in figure 9.29 on page 371.

Delocalization allows the pi electrons to move about the molecule, leading to a more stable (lower energy) state.

## VII. Molecular Orbitals

### A. Molecular Orbital Theory

Molecular orbital theory assumes the formation of molecular orbitals from atomic orbitals. M.O. theory uses all of the electrons in the molecule, NOT JUST THE VALENCE ELECTRONS. However, in writing the molecular orbital electron configuration, just the valence electrons can be shown. The molecular orbitals formed are sigma and pi bonding and antibonding orbitals.

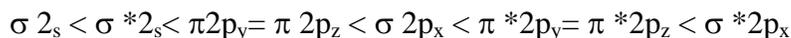
### B. Bond Order

The bond order is calculated as follows:

Bond Order =  $\frac{1}{2}$  (# of electrons in bonding M.O.'s - # electrons in antibonding M.O.'s)

## VIII. Second-Row Diatomic Molecules

The energy of molecular orbitals for the second row diatomic molecules is shown in Figure 9.48 on page 384 of the text. The order of electron filling of the valence electrons is as shown below:



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**Chemistry 1201 Review/Preview  
Chapter Ten Review Guide**

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**Spring 2007**

**Gases**

**I. Characteristics of Gases**

Gases have many properties that are quite different from those of liquids and solids, as listed below.

1. Gases expand to fill their container.
2. The volume of a gas decreases when pressure is applied, and increases when pressure is decreased.
3. Gases mix freely to form homogeneous mixtures (solutions) regardless of the type of gas or the relative amounts of each gas.
4. Model that describes a gas as a collection of mo.
5. The individual molecules in a gas are very far apart. Most of the volume of a gas is empty space.
6. The molecules of a gas are spaced far apart and move around rapidly and randomly, colliding occasionally with one another or with the walls of the container.
7. Intermolecular forces are essentially non-existent because the molecules are too far apart and moving too rapidly to attract or repel each other.

Gas samples are typically described in terms of four variables -- pressure, volume, temperature, and number of moles present.

**II. Pressure**

**A. Pressure definition and units**

Pressure is defined as force per unit area ( $P = F/A$ ). The common units for pressure are:

mm Hg (also known as torr), atmospheres, pascals, and kilopascals. See Section 10.1 of the text for conversion factors for pressure units.

Do Sample Exercises 10.1 and 10.2 (without looking at the solutions.)

**B. Instruments for measuring pressure**

1. barometer - used for measuring atmospheric pressure
2. manometer - used for measuring the pressure of a gas within a closed container.

(Study Figures 10.2 and 10.3 on pages 402 and 403 of the text to fully understand how barometers and manometers work.)

**III. The Gas Laws**

**A. The Pressure-Volume Relationship: Boyle's Law**

At a fixed temperature, the volume of a gas is inversely proportional to its pressure. Mathematically stated, Boyle's Law is:  $P_1V_1 = P_2V_2$ .

**B. The Temperature-Volume Relationship: Charles Law**

At a fixed pressure, the volume of a gas is directly proportional to its absolute temperature. Mathematically stated:  $V_1/T_1 = V_2/T_2$

**C. The Quantity-Volume Relationship: Avogadro's Law:**

The volume of a gas kept at a constant temperature and pressure is directly proportional to the number of moles of the gas. Mathematically stated:  $V = \text{constant} \times n$ , where  $n$  represents the number of moles of gas.

(Note that one mole of any gas at STP occupies 22.4 liters.)

**D. The Combined Gas Law:**

When Boyle's Law, Charles Law and Avogadro's Law are combined, the combined gas law:  $P_1V_1/T_1n_1 = P_2V_2/T_2n_2$

**IV. The Ideal-Gas Equation**

The ideal-gas equation is  $PV = nRT$ , where P is in atmospheres, V is in liters, T is in kelvins, and the value of R = 0.08206 liter·atm/mole·kelvin.

You should notice that the ideal gas law contains five terms:

P, V, n, R, and T. Four of the terms are variables (P, V, n, and T) and one is a constant, R. The value of R is  $0.0821 \frac{\text{L} \cdot \text{atm}}{\text{K} \cdot \text{mole}}$

Problems involving the ideal gas law will generally specify three of the four variables, and you will be asked to solve for the fourth one. When using the ideal gas law you should always make sure that the variables are expressed in the units that are compatible with the units of R, the ideal gas constant. If they are not in the correct units, you must convert them to the correct units before substituting them into the equation.

**V. Further Applications of the Ideal-Gas Equation**

**A. Gas Density Calculations**

The ideal gas law can be rearranged to derive an expression that involves density. (Remember density units are mass/volume.)

$d = PM/RT$ , where M is the molar mass of the gas

**B. Calculating the molar mass from gas density data can be accomplished by using the relationship:  $M = dRT/P$**

**C. Volumes of Gases in Chemical Reactions**

At STP, one mole of any gas occupies 22.4 liters/mole. Additionally the moles of any gas at a different temperature and/or pressure can be calculated using the ideal gas law. After the moles have been determined, gas law stoichiometry problems are like all other stoichiometry problems. (See pages 415-416 and do Sample Exercise 10.9 for practice with these types of problems.)

**VI. Gas Mixtures and Partial Pressures**

**A. Partial Pressures and Mole Fractions**

In a mixture containing gases 1 & 2, the total pressure  $P_t = P_1 + P_2$

Additionally,  $P_1 = X_1P_t$  and  $P_2 = X_2P_t$ , where X represents the mole fraction.

**B. Collecting Gases Over Water**

Whenever a gas is collected over water, the pressure of the gas in question must be obtained by subtracting the vapor pressure of water at that temperature.

**VII. Kinetic-Molecular Theory of Gases**

The essential statements of the kinetic theory are found on page 421 in the text. Pay particular attention to the equation for the root mean square speed of a gas particle shown on page 421 of the text.

**A. Application to Boyles Law**

When there is a volume increase at constant temperature, the molecules have to move farther before hitting the walls of the container. There are therefore fewer collisions and the pressure decreases.

**B. Effect of temperature increase at constant volume**

When the temperature is increased at constant volume, the molecules move faster, hitting the walls of the container harder, thus increasing the pressure exerted by the gas.

## VIII. Molecular Effusion and Diffusion

- A. **Effusion** – the rate at which a gas escapes through a tiny hole into a vacuum
- B. **Diffusion** – the rate at which a gas spreads throughout space or throughout another gas.
- C. **Calculating Rates of Effusion and Diffusion**

Under the same conditions of temperature and pressure, the rates at which gases diffuse or effuse are inversely proportional to the square roots of their molar masses.

See formulas and examples on pages 424 – 425 of the textbook.

Note that rates and times are inversely proportional to each other!

- D. **Diffusion and Mean Free Path**

The mean free path is the distance a gas molecule travels before it collides with another molecule. The mean free path decreases as the pressure increases.

## IX. Real Gases: Deviations From Ideal Behavior

Since ideality assumes that there are no interactions between the molecules of a gas, and that volume of the gas particles is negligibly small, assumptions based on ideality are only as good as the assumptions are true. For real gas, there are some interactions between particles, and the molecules do occupy a finite volume. Gases deviate most from ideality at high pressure and low temperature.

When corrections for non-ideal behavior are introduced into the ideal gas equation, the relationship becomes the van der Waals equation, found on page 429 of the textbook. (Study this equation so that you know what the corrections are and why they are made.)

**NOTE: This review sheet is for Chem 1201 Review/Preview attendees and is to be used in conjunction with lecture notes, handouts, problem sets, the textbook, and any other material from Chemistry 1201.**

**Chemistry 1201 Review/Preview  
Chapter Eleven Review Guide**

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**Spring 2007**

**Intermolecular Forces, Liquids, and Solids**

**I. A Molecular Comparison of Gases, Liquids, and Solids**

Gases, liquids, and solids are very different in the way the molecules move in the substance. As we saw in Chapter 10, the molecules of gases move freely to occupy the entire volume of the container. The molecules of liquids, however, are held close together, but move freely enough to assume the shape of the container the liquid is in. The molecules of a solid are held close together, and are essentially locked into place so that the shape of a solid does not change.

The three states (or phases) -- gas, liquid, and solid -- can be changed into one another by heating, cooling, or changing the pressure, as described in detail in Section 11.1 of the textbook.

**II. Intermolecular Forces**

**A. Description of Intermolecular Forces**

The attraction *between* the molecules in a substance is referred to as intermolecular forces. Note that intermolecular forces are different from chemical bonds within a molecule. (Bonds are *intramolecular* forces.) Intermolecular forces cause molecules to stick to one another to form a liquid or a solid under normal conditions. Substances with very weak intermolecular forces are likely to be gases at room temperature, those with very strong intermolecular forces are likely to be solids, and those that have intermolecular forces between very weak and very strong are likely to be liquids at room temperature. (See Section 11.2 on page 445 for more discussion of this.)

**B. Types of Intermolecular Forces**

When an ionic compound is dissolved in a polar solvent, the intermolecular forces are referred to as ion-dipole forces

There are three types of intermolecular forces that operate between neutral molecules. These are listed below:

1. Dipole - dipole forces  
These forces are present in neutral, polar molecules. The partially positive end of one molecule is attracted to the partially negative end of another molecule. (See Figure 11.4 in the textbook.)
2. London Dispersion Forces (also called London forces or dispersion forces) - These forces are present in nonpolar substances. The magnitude of London dispersion forces depends on the polarizability of the molecule and the shape of the molecule. *Polarizability* refers to the ease of distortion of the electron cloud associated with an atom or molecule. Polarizability generally increases with increasing number of electrons and larger size of the atom or molecule. (See Table 10.3 of the text.) With regard to shape, molecules with shapes that result in more molecular surface area give rise to higher dispersion forces than shapes with less surface area. Therefore, linear

molecules have larger London dispersion forces than spherical molecules. Note that all *molecules* have London dispersion forces because all molecules contain electron clouds.

3. **Hydrogen Bonding** - a special type of very strong dipole-dipole interaction, present in compounds that contain an O-H, N-H, or H-F bond. No compounds can exhibit hydrogen bonding without one of these three bonding types.

### C. **Comparing Intermolecular Forces**

The order of strength of intermolecular forces is as follows:

*H-bonding > dipole-dipole interactions > dispersion forces*

All molecules (polar and non-polar) experience London dispersion forces; polar molecules experience London dispersion forces and dipole-dipole interactions; molecules with an HF, NH, or OH bond experience London dispersion forces, dipole-dipole interactions, and hydrogen bonding.

NOTE: See Figure 11.12 on page 452 for a very important flow chart for determining intermolecular forces. It is crucial to understand this chart. Do not memorize it, understand and be able to apply it.

### D. **Relationship of Intermolecular Forces and Boiling Point**

The stronger the intermolecular attractive forces between the molecules in a substance, the higher the boiling point. (See Tables 11.2 and 11.3 in the text.)

## III. **Some Properties of Liquids**

- A. **Viscosity - a fluid's resistance to flow.** The stronger the intermolecular forces, the higher the viscosity. Viscosity decreases with an increase in temperature and increases with a decrease in temperature. For example, the thickness of oil decreases as the temperature is increased.
- B. **Surface Tension** - the resistance of a liquid to spreading out and increasing its surface area. Liquids with stronger intermolecular forces generally have higher surface tension.

## IV. **Phase Changes**

### A. **Definition**

Phase Changes are transformations that substances undergo as they physically change state, but maintain their chemical identity, e.g. solid to liquid, liquid to gas, etc. (See section 11.4 of the text and lecture notes for a complete discussion of the relationship of enthalpy, entropy, and free energy to phase changes. Pay particular attention to the meaning of the terms *heat of fusion* and *heat of vaporization*.)

### B. **Heating and Cooling Curves -**

Heating and cooling curves graph the relationship between the temperature of a substance and the amount of heat added. (See figure 11.19 in the text for a description of the heating curve for water.)

### C. **Critical Temperature and Pressure**

The critical temperature is the highest temperature at which a gas can be converted to its liquid phase. The critical pressure is the pressure of the gas at this temperature.

(See Table 11.5 on page 458 of the textbook for the values of the critical temperatures and pressures of some common substances.)

## V. Vapor pressure

### A. Definition

Vapor pressure is the pressure exerted by the molecules of a gas in contact with its liquid. Equilibrium vapor pressure is the pressure measured under conditions of dynamic equilibrium between the liquid and the vapor. The rates of evaporation and condensation are equal at this point of dynamic equilibrium. Generally, the term vapor pressure is used to denote equilibrium vapor pressure. The equilibrium vapor pressure rises with temperature. See Figure 11.24 in the text to see how the vapor pressure varies with a rise in temperature. Substances with a high vapor pressure (evaporate readily) are said to be *volatile*.

### B. Vapor Pressure and Boiling Point

The boiling point of a liquid is the temperature at which the vapor pressure is equal to the external pressure. The normal boiling point of a liquid is the temperature at which the vapor pressure is equal to 760 mm Hg.

## VI. Phase Diagrams

A phase diagram summarizes the conditions at which a particular substance exists as a solid, liquid, or gas. See figure 11.27 and in the text for a comparison of the phase diagrams for water and carbon dioxide.

## VII. Structures of Solids

### A. Types of solids -

1. Crystalline solids - have long-range order; highly regular arrangement of components.
2. Amorphous solids - no long range order; have considerable disorder in their structures.

### B. Unit Cells

1. Unit cell - the smallest repeating unit of a crystal.
2. Three Types of Cubic Unit Cells - primitive cubic, body centered cubic, face centered cubic. (See Figure 11.33 for the pictures of these unit cells.)
3. The atoms per unit cell in each type of unit cell are given below.
  - a. Primitive Cubic - 1 atom/unit cell (8 corner atoms  $\times$   $1/8$ )
  - b. Face-Centered Cubic - 4 atoms per unit cell (8 corner atoms  $\times$   $1/8$  + 6 face atoms  $\times$   $1/2$ )
  - c. Body-Centered Cubic - 2 atoms/unit cell (8 corner atoms  $\times$   $1/8$  + 1 central atom  $\times$  1)

### C. Close Packing of Spheres

1. Hexagonal Closest Packing - noncubic unit cell, 74% efficiency
2. Cubic Closest Packing - face centered cubic unit cell, 74% efficiency (See page 469 of the text for a complete description of the kinds of spherical packing.)

## VIII. Bonding in Solids

Solids are classified based on the types of forces holding the particles together. The classes of solids are described below.

- ### A. Molecular Solids
- consist of atoms or molecules held together by intermolecular forces. Examples are ice ( $\text{H}_2\text{O}$ ) and dry ice ( $\text{CO}_2$ )

- B. Covalent-Network Solids** – consist of atoms held together by covalent bonds. Examples are diamond (carbon) and quartz ( $\text{SiO}_2$ )
- C. Ionic Solids** – consist of ions held together by ionic bonds. Examples are  $\text{NaCl}$ ,  $\text{CaBr}_2$ , and  $\text{MgO}$
- D. Metallic Solids** – consist of metal atoms held together by valence electrons delocalized around metal cations throughout the entire solid. (See Figure 11.45 for a picture of what bonding in metals looks like.)

**NOTE: This review sheet is for Chem 1201 Review/Preview attendees and is to be used in conjunction with lecture notes, handouts, problem sets, the textbook, and any other material from Chemistry 1201.**

**Chemistry 1201 Review/Preview  
Chapter Thirteen Review Guide**

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**Properties of Solutions**

**I. The Solution Process**

**A. Terms and Concepts Defining Solutions**

1. Solutions - homogeneous mixtures
2. Solute - the substance that is dissolved (present in the smaller amount).
3. Solvent - the dissolving substance (present in the larger amount).
4. Solvation – the interactions between the solute and solvent
5. Hydration – the interactions between a solute and water

**B. Types of Solutions**

- |                  |                   |                    |
|------------------|-------------------|--------------------|
| 1. Solid – Solid | 2. Solid – Liquid | 3. Liquid - Liquid |
| 4. Liquid – Gas  | 5. Solid – Gas    | 7. Gas - Gas       |

**C. Energy Change and the Solution Process**

The overall enthalpy change when forming a solution,  $\Delta H_{\text{soln}}$ , involves three steps:

1. Separation of the solute molecules
2. Separation of the solvent molecules
3. Formation of interactions between the solute and the solvent

The overall enthalpy is calculated by adding the enthalpy changes for the three steps:  $\Delta H_{\text{soln}} = \Delta H_1 + \Delta H_2 + \Delta H_3$

$\Delta H_1$  and  $\Delta H_2$  are always positive values, and  $\Delta H_3$  is always a negative value.

The heat of solution can be positive or negative, depending on the size of the individual  $\Delta H$  values. The sizes of the individual  $\Delta H$  values depend on the strength of the solute-solute, solvent-solvent, and solute-solvent intermolecular attractive forces.

**D. Solution Formation, Spontaneity, and Disorder**

*Entropy* is the amount of disorder in system. Changes that result in an increase in entropy tend to occur spontaneously. Because the formation of solutions results in an increase in entropy, solutions will form spontaneously unless the solute-solute and solvent-solvent interactions are too strong.

**II. Saturated Solutions and Solubility**

When a solute dissolves in a solvent, a solution is formed. When the solvent is removed from the solution, crystallization occurs.

**A. Saturated solution** – a solution that contains the maximum amount of solute possible at a given temperature. (The solution contains undissolved solute in equilibrium with the solution.)

**B. Unsaturated solution** – a solution that contains less solute than a saturated solution.

- C. **Supersaturated solution** – a solution that contains more solute than is necessary to form a saturated solution. Supersaturated solutions are formed by making a saturated solution and then slowly cooling the solution to a lower temperature. (See Figure 13.1 on page 536 for a picture of a supersaturated solution of sodium acetate.)

### III. Factors Affecting Solubility

#### A. Solute – Solvent Interactions

General rule of solubility - "like dissolves like" (Substances with similar intermolecular attractive forces will dissolve in each other.) This is true because the intermolecular forces between solute and solvent particles will be strong enough to overcome the solute-solute and solvent-solvent interactions. Substances that dissolve completely in each other are said to be *miscible*; substances that do not mix are *immiscible*. (See Table 13.3 for data illustrating the "like dissolves like" rule.)

- B. **Pressure** - For solutions with a gaseous solute, the solubility of the gas increases with an increase in the partial pressure of the gas. This relationship is described quantitatively by Henry's Law:  $S_g = k \cdot P_g$  where  $S_g$  is the concentration of the dissolved gas in moles/liter,  $P_g$  = partial pressure of gaseous solute above the solution, and  $k$  is the Henry's law constant.

(Do Sample Exercise 13.3 for practice with Henry's Law calculations.)

- C. **Temperature** - For most solutions of a solid in a liquid (although there are a few exceptions), the solubility increases with an increase in temperature. For solutions with a gaseous solute, the solubility decreases with an increase in temperature.

### IV. Ways of Expressing Concentration

#### A. Mass Percentage, ppm, and ppb

1. Mass Percent - grams of solute per hundred grams of solution

$$\text{Mass \% of component} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$$

2. Parts Per Million (ppm) - grams of solute per million grams of solution

$$\text{ppm} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 10^6$$

3. Parts Per Billion (ppb) - grams of solute per billion grams of solution

$$\text{ppb} = \frac{\text{grams of solute}}{\text{grams of solution}} \times 10^9$$

#### B. Mole Fraction, Molarity, and Molality

1. Mole fraction - the mole fraction of any solution component

$$\text{Mole fraction of component} = \frac{\text{moles of that component}}{\text{total moles of solution}}$$

2. Molarity (M) - moles of solute per liter of solution

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

3. Molality (m) - moles of solute per kilogram of solvent

$$\text{Molality} = \frac{\text{moles of solute}}{\text{Kilograms of solvent}}$$

(Do practice exercises 13.4 and 13.5 for practice in calculating concentrations.)

**C. Conversion of Concentration Units**

Because the concentration units contain grams, moles, mole fraction, volume of solution, mass of solvent, and other quantities that can be calculated from each other, concentration units can be interconverted. Carefully study pages 546 and 547 of the text and Figure 13.19 for practice in converting between concentration units.

**VI. Colligative Properties** - properties that depend only on the number of solute particles present and not on the nature of the particles.

**A. Vapor Pressure Lowering** -

1. Non-volatile Solutes - The presence of a non-volatile solute in a solvent results in a lower vapor pressure for the solution than for the pure solvent. The extent to which the vapor pressure is lowered depends on the mole fraction of the solute (and thus solvent) in the solution. Mathematically speaking this is expressed by Raoult's Law:  $P_A = X_A \times P_A^0$  where  $P_A$  is the vapor pressure above the solution,  $P_A^0$  is the vapor pressure of the pure solvent, and  $X_A$  is the mole fraction of solvent in the solution.

**B. Boiling Point Elevation & Freezing Point Depression of Solutions** -

The boiling point of a solution is higher than that of the pure solvent, and the freezing point of a solution is lower than that of the pure solvent. The mathematical relationships that express these phenomena are shown below:

$$\Delta T_b = K_b \cdot \text{molality} \quad \Delta T_f = K_f \cdot \text{molality}$$

where  $K_b$  and  $K_f$  are constants that depend on the solvent only

**C. Osmosis** - The pressure which must be applied to a solution that is in contact with a pure solvent (via a semi-permeable membrane) in order for there to be no net movement of solvent particles across the membrane. The magnitude of the osmotic pressure,  $\Pi$ , depends on the molarity of the solution according to the following equation:

$$\Pi = \text{Molarity} \cdot R \cdot T \quad (\text{where } R \text{ is the gas constant})$$

**D. Determination of Molar Mass from Colligative Properties**

The colligative property is used to determine the number of moles of unknown substance,  $X$ .

$$\text{Molar mass} = \text{grams of } X / \text{moles of } X$$

The grams of  $X$  are usually given in the problem. Moles of  $X$  can be calculated from the molality or the molarity. (Do Sample Exercise 13.13.)

**VI. Colloids** – mixtures that contain large particles suspended in a solvent.

**A. The Tyndall Effect** – the scattering of light by particles in a colloid

**B. Hydrophilic and Hydrophobic Colloids**

Hydrophilic colloids contain particles that are soluble in water; hydrophobic colloids contain particles that are not soluble in water. Hydrophobic colloids must be stabilized if the particles are to remain suspended in the solvent.

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