**THINGS TO KNOW FOR DR. KREISER’S CHEM 1030 CLASS:**

1. Show up for lecture and take good notes. If you can, record the lecture and review it for parts you missed.
2. If you miss a class, see someone for the notes! Review the lesson in the book and visit Dr. Kreiser during office hours if you have questions or email the class tutor (kjmanchester@ccri.edu).
3. Memorize your ions and their respective charges! Know them as well as you know your own name!
4. READ THE TEXT! This isn’t on the syllabus because it is assumed!
5. Know how to read a periodic chart

As you learn new things, remember them! Things like…

The soluble and insoluble compounds chart on chapter 4 of Chang (precipitation reactions)

The steps to writing a balanced net ionic equation (also in Chapter 4)

The equation formulas used for density, specific heat, gas laws, etc. – you will not be given these on tests and quizzes!

And finally…

Have fun…it makes the learning easier!

STUDY TIPS FOR TEXTBOOK READING

1. Read the chapter summaries FIRST, then read the chapter a little bit at a time
2. Try to complete the practice problems in the chapter, using the examples as a step-by-step guide
3. If you don’t understand the problem or a paragraph, break it down by sentences and use the glossary to look up unfamiliar words
4. If the professor says you don’t need to read something don’t *feel that you need to read it anyway! You will only be creating more work for yourself!*
5. Keep your lecture notes by your side as you read and when you find a “match” write the page number in your notes.

**DIMENSIONAL ANALYSIS**

When doing a conversion, dimensional analysis is just one big fraction that is equal to “1” in all places EXCEPT for the first spot, which is where you place the unit you are converting.

Example: How many millimeters are in 5.0 kilometers?

5.0 km x (1,000 m/1 km) x (100 cm/1m) x (10 mm/1 cm) Do the math…

The answer is 5,000,000 mm or 5.0 x 106 mm

When using dimensional analysis to balance chemical equations, the same premise applies; you will learn more about this in the next tutorial packet.

**SIGNIFICANT FIGURES**

“Sig Figs” are the numbers that survive the conversion to scientific notation

Ex.: .00005 has ONE significant figure; all of the zeros are placeholders to show the size of the number

 .000050 has TWO significant figures; the zero that follows the five shows precision

When adding, you can GAIN sig figs: 9.0 + 1.0 = 10.0

When subtracting you can lose sig figs: 11.0 – 10.0 = 1.0

**NOTE: Unless you are truncating your answer, you do not gain or lose sig figs after the decimal point when adding! (When subtracting, you can lose them only if your answer is less than .1).**

When using sig figs do the math as normal and truncate your answer to match the least number of significant figures in your information!

Example:

 232.4 If I truncate the numbers too soon, my answer is incorrect!

 155.10 Add all the numbers present, and then truncate the final answer!

 333.751

+836.5437

1,557.7947 is the exact answer; to the proper number of sig figs the answer is 1,158 or 1.158 x 103.

If I truncated the numbers before adding my answer would be 1556 or 1.556 x 103. It would also be wrong!

**HOW TO DO TEMPERATURE CONVERSIONS**

To convert between Fahrenheit and Celsius

From °C to °F: (9/5 x °C) + 32°

From °F to °C: (°F - 32°) x 5/9

If your calculator does not accept fractions, get one that does! In the meantime, use the decimal equivalents: 1.8 = 9/5 and .556 = 5/9

To convert between Celsius and Kelvin

To K from °C: ADD 273.15

From K to °C: SUBTRACT 273.15

Example: 0 °C = 273.15 K (0 + 273.15)

0 K = -273.15 °C (0 – 273.15)

If you ever confuse your conversions, try converting from 0 to remember if you have to add or subtract

**The Density Formula**

Density is an intrinsic quality, meaning you can change the mass or volume and the density will change according to scale. The formula for density is:

Density = Mass/Volume or D = M/V

This formula can be rearranged to find mass or volume by applying simple algebra.

If D = M/V, then M = D x V

If D = M/V, then V = M/D

An easy way to remember these formulas is with a “wedge” chart:

Remove the letter you need to solve for;

M

D

V

Do the operation that remains:

If the letters are side by side, multiply; if the letters are up and down, divide.

**Specific Heat**

Heat is not the same thing as temperature. Heat is a form of energy and is measured in joules (J) or kilojoules (kJ).Temperature is a measurement of how intense an effect heat has on something.

 **Specific heat** is defined as the amount of energy (in joules or kilojoules) that it takes to raise 1 gram of a substance 1 degree Celsius. (In Biology, that substance is water and we call it a calorie!).

All things contain energy and that energy cannot be created or destroyed – it is merely transferred as heat. In Chemistry, we measure that heat transfer with the following equation:

q = kmΔt

q

k

m

Δt

“q” = heat (in joules or kilojoules)

“k” = the specific heat

“Δt” = the change in temperature (in degrees Celsius)

**For directions on how to use chart,**

**see previous page!**

**PROTONS, NEUTRONS, AND ELECTRONS (OH MY!)**

An atom is the smallest amount of an element you can have and still have that element.

Atoms are made up of sub-atomic particles called protons, electrons, and neutrons

A proton has a positive charge

An electron has a negative charge

A neutron has no charge – it is neutral

A proton and a neutron both weight 1 atomic mass unit (amu), which is equal to 1 gram; an electron has negligible weight – it only weighs 9.1 x10-31 kg! (That’s 9.1 x 10-33 grams – that’s 31 zeros followed by a 91)

An atoms weight is based upon how many protons and neutrons it contains

An atom’s atomic number is the number of protons it contains

 An atom in its elemental form has the same number of electrons as it does protons

 An atom that has been ionized will have more of one than the other

 An atom with more neutrons than protons is called an**isotope**.

**ATOMIC STRUCTURE AND IONIZATION**

An easy way to visualize atomic structure is to think of an atom’s structure as a wedding reception:

The bride and groom’s “sweetheart table” holds ONLY two people; an atom’s first valence (aka “shell”) holds ONLY two electrons

All subsequent tables at the wedding can seat UP TO eight guests each; all subsequent atomic valences can hold UP TO eight electrons each

A table for eight with only one person sitting at it can be combined with a table for eight where only seven people are seated (there is room for one more); an atom with only one electron in its outermost valence can combine with an atom that has seven electrons in its outmost valence.

The atoms that give up or take on electron(s) have been **ionized*.*** These atoms no longer have an equal ratio of protons to electrons and are no longer electrically neutral.

An atom that has LOST an electron now has more protons (positively charged particles) and will have a **positive** charge. The positive charge is equal to the number of electrons lost.

An atom that has GAINED an electron now has more electrons (negatively charged particles) and will have a **negative** charge. The negative charge is equal to the number of electrons gained.

**HOW TO READ A PERIODIC CHART**

A periodic chart can tell you a lot about an element – especially how it ionizes! Group B elements are known as transition metals, and can have more than one ionic charge. The reason for this is more than this class requires you to know, so don’t worry about *why*; just memorize the charges on your ion sheet!

Group 1A elements all have 1 electron (written e-) in their outermost valence. They have an ionic charge of 1+

Group 2A elements all have 2 e- in their outermost valence. They have an ionic charge of 2+

Group 3A elements all have 3 e- in their outermost valence. They have an ionic charge of 3+

**DO YOU SEE A PATTERN DEVELOPING HERE?**

Group 4A elements all have 4 e- in their outermost valence. They do not ionize but rather share electrons in something called a **covalent bond**.

Group 5A elements all have 5 e- in their outermost valence. They have an ionic charge of 5+ or 3-. Since it requires less energy to take on 3 e- than it does to give away 5 e- the charge on this ion is generally 3-

Group 6A elements all have 6 e- in their outermost valence. They have an ionic charge of 6+ or 2-. Since it requires less energy to take on 2 e- than it does to give away 6 e- the charge on this ion is generally 2-

Group 7A elements all have 7 e- in their outermost valence. They have an ionic charge of 7+ or 1-. Since it requires less energy to take on 1 e- than it does to give away 7 e- the charge on this ion is generally 1-

Group 8A elements all have 8 e- in their outermost valence. They are electronically stable and therefore do not react. They are known as the Noble gases and are considered to be chemically **inert**.

**WHY ARE ION CHARGES SO IMPORTANT?**

Elements combine to form compounds because they seek to become electronically stable/neutral by filling their outermost valence. Some elements do this by giving up the e- in their outermost valence, thus emptying it and making the one beneath it the outermost (and full) valence; others will take on electrons in order to create a complete component of eight e- in their outermost valence.

Compounds must be electrically neutral – this means that their ionic charges have to add up to zero.

Ex.: Na1+ + Cl1-‑ 🡪 NaCl

Notice that the +1 and the -1 add up to zero! Not all compounds balance so easily.

Ex.: Mg2+ + Br1- 🡪 MgBr

Notice that this is not electronically neutral, as written: +2 + (-1) ≠0

The correct formula is Mg2+ + 2Br1- 🡪 MgBr2

If you do not know your ion charges everything becomes guesswork – and the law of averages says that you will usually guess wrong!

A positively charged ion is called a **cation**

A negatively charged ion is called an **anion**

**Naming Chemical Compounds**

There are three ways to name compounds – the “-ous/-ic” system and the Stock system are used for ionic compounds while the covalent system is used for compounds that share electrons.

**The –ous/-ic System**

When using the –ous/-ic system, you must change the ending of the cation to reflect the charge

(Hint: -ous is *usually* the smaller charge).

With this system, the cation and the anion keep their names and you list them in order, cation first, then anion.

Examples: sodium chloride (NaCl); potassium sulfate (K2SO4); aluminum bromide (AlBr3)

If a compound contains a metal with more than one charge, look to the anion to see what the charge on the cation must be in order to keep the compound electrically neutral. (Note: In a situation such as this, the compound will be balanced for you on a test or a quiz).

Example: FeSO4 is *ferrous* sulfate; Fe2(SO4)3 is *ferric* sulfate. Because the charge on the iron (Fe) can be +2 or +3 we need to look at the charge on the anion to figure the charge on the cation. (SO42- x 3 = -6; therefore, Fe2 must be 6+. 6+/2 = 3+ or ferric).

In the case of an ion having more than one charge (and it will generally be the cation; anions are the exception, not the rule), you can use the Stock system to name the compound.

**The Stock System**

The Stock system uses Roman numerals to describe the charge on the cation. (The exception to this is mercurous and mercuric; in this case, the Roman numeral refers to the number of atoms that make up the cation).

The Stock system is used *only* when the cation can have more than one charge or (as in the case of mercurous) is made up of multiple atoms of the same element.

Just like the –ous/-ic system, the Stock system looks to the anion to determine the charge on the cation.

Example: arsenic chloride is written AsCl5­. Because chloride has a charge of 1- and there are 5 chloride ions in the compound arsenic must have a charge of 5+ for the compound to be electrically neutral. Therefore, the compound name in the Stock system is **arsenic (V) chloride**.

Another example would be the one from the previous page – ferrous sulfate (FeSO4) and ferric sulfate [Fe2(SO4)2]. Ferrous sulfate becomes iron (II) sulfate and ferric sulfate becomes iron (III) sulfate.

The “-ous/-ic” system and the Stock system are used for naming ionic compounds. A third system, the covalent system, is used for naming compounds that share elections (rather than give or receive electrons, like ionic compounds).

**The Covalent System**

The covalent system uses prefixes to describe how many atoms of each element are present in a compound. The prefixes are attached to the element name. The prefixes are:

mono – one hexa – six **Note that two is “di”, not “bi”; in Chemistry “bi” is the prefix**

di – two septa – seven **that means hydrogen is present in an ion.**

tri – three octa - eight

tetra – four nona – nine

penta – 5 deca – ten

Under the covalent system the chemical name for water – H­2O – is dihydrogen monoxide. (Google that name for some Chemistry fun!).

Other examples of covalent chemical names are carbon dioxide (CO2); sulfur hexafluoride (SF6); and dinitrogen tetroxide (N2O4).

**HOW TO BALANCE A CHEMICAL EQUATION**

1. KNOW YOUR ION CHARGES!! I will remind you again that if you do not know these, you cannot correctly balance a chemical equation!
2. All compounds are electrically neutral, so you need the positive charge (cation) to equal the negative charge (anion).
3. If the cation and the anion have unequal charges, you must find their lowest common multiple (LCM)
	1. Example: Mg2+ + PO43- 🡪 ???

The LCM of 2 and 3 is 6

1. Multiply the charge by “x” or “y” to find the number of atoms/moles needed to react
	1. Example: “x” Mg2+ + “y” PO43- 🡪 ???

(2+)(x) = 6 (3-)(y) = 6 x = 3; y = 2

1. Add the x and y numbers to the reactant (left side) of the equation as *coefficients*; add them to the right side as *subscripts*, making sure that they are written in the lowest ratio possible (ex. 3:2 not 6:4).

**The answer to this equation is 3Mg + 2PO4 🡪 Mg3(PO4)2**

In words, this equation reads “Three moles of magnesium combine with two moles of phosphate to form one mole of magnesium phosphate”

Because PO4 is a polyatomic ion we use parentheses to show that the entire ion is “taken twice”

*You cannot divide a polyatomic ion into its parts – it remains as written!*

**HOW TO FIGURE AN EMPIRICAL FORMULA (KNOWN MOLECULAR MASS)**

An **empirical formula** is the molar ratio – *not the percentage ratio –* of elements in a compound.

How to find it is best taught by example:

A compound has a molar mass of 160.0 g/mol. A representative sample (of less than 1 mole!) was analyzed and found to be 50.00 grams copper, 30.52 grams sulfur, and 60.52 grams oxygen. What is the empirical formula?

1. Write the grams per mole of each element in the compound:

Cu = 63.55 g/mol S = 32.07 g/mol O = 16.00 g/mol

1. Convert the grams of each element in the compound to moles of each element by dividing by the respective molar mass. This will give you the molar ratio of each element in the compound. (Note: Round up or down to the nearest whole number).

50.00 g Cu/63.55 g/mol = .7868 mols ≈ 1 mol

30.52 g S/32.07 g/mol = .9517 mols ≈ 1 mol

60.52 g O/16.00 g/mol = 3.783 mols ≈ 4 mols

1. The ratio of elements in the compound is 1 Cu: 1 S: 4 O; the empirical formula is CuSO4.

If you do not know the molecular mass of the compound, see the next page for instructions on how to find the empirical formula based upon elemental percentage!

**HOW TO FIGURE AN EMPIRICAL FORMULA (UNKNOWN MOLECULAR MASS)**

1. Assume 100 grams of a compound. Memorize this step because if you forget it you will not be able to solve the problem before you!

2. Take the percentage of each element in the compound and change it to grams:

 Example: 33.33% = 33.33 grams

 42.25% = 42.25 grams

 20.42% = 20.42 grams

NOTE THAT IT ADDS UP TO 100 GRAMS!

3. MOST IMPORTANT! For each element in the compound you must convert from grams to moles!

 Grams of an element/molecular mass = number of moles

4. Take the smallest number of moles in the compound and divide all other elements in the compound (now in moles) by that number!

EXAMPLE ON THE NEXT PAGE!

Example: A compound contains 66.0% oxygen and 34.0% nitrogen. What is the empirical formula, rounded up/down into whole numbers?

Step 1: Assume 100 grams of the compound. 100% = 100 grams

Step 2: Change percentages to grams 66.0% O = 66.0 g O O = 16.00 g/mol

34.0% N = 34.0 g N N = 14.01 g/mol

Step 3: Calculate the number of moles of 66.0 g O/16.00 g/mol = 4.125 mols

each element in the compound based upon 34.0 g N/14.01 g/mol = 2.427 mols

your answers from step 2.

Step 4: Take the smallest number of moles 2.427 mols N/2.427 mols = 1 mol N

in the compound and divide all other elements 4.125 mols O/2.427 mols = 1.70 mols O

in the compound (now in moles) by that number!

Step 5: Round up each answer to the nearest 1 mol N:1.70 mols O ≈ 1N:2 O

whole number to get your answer **The empirical formula is NO2**

**WHAT IS A LIMITING REAGENT?**

What is a limiting reagent? Simply put, it is the chemical that will run out first when you are creating a compound.

For example, if I have a cake mix that needs 1 box of mix, 3 eggs and a cup of milk to complete and I have 2 cake mixes, 3 cups of milk, and a dozen eggs how many cakes can I make?

1 cake = 3 eggs. 1 cup milk, 1 mix

2 cakes = 6 eggs, 2 cups milk, 2 mix

3 cakes = 9 eggs, 3 cups milk, 3 mix

4 cakes = 12 eggs, 4 cups milk, 4 mix

Although I have milk enough for three cakes; and enough eggs for four cakes; I only have two cake mixes, so the mix is my **limiting reagent**, *the element I will* ***run out of first*** *and that will determine how many cakes I can make*. Substitute a chemical formula for a cake recipe and you get the picture!

**HOW TO FIGURE LIMITING REAGENTS**

This is another problem that is best shown by example. See the next page for an example.

Example: You have 10.0 grams of magnesium and 20.0 grams of chloride. How many moles of magnesium chloride (MgCl2) can you make?

1. Write and balance the equation, which is the “recipe” for this compound: Mg + 2 Cl 🡪 MgCl2
2. Look at the ratio between the two elements/ions: 1 Mg + 2Cl = 1:2 ratio Mg:Cl
3. Use dimensional analysis to figure out which one is the limiting reagent:

10 g Mg x (1mol Mg/24.31 g Mg) x (2 mols Cl/1 mol Mg) x (35.45 g Cl/1 mol Cl) = 29.16 grams Cl

This is how many grams of chloride are needed to react all 10.0 grams of magnesium. Since you only have 20.0 grams of chloride, this is the limiting reagent.

**BUT WAIT! YOU ARE NOT DONE YET!!!!!!!!!!!!!!**

**HOW TO FIGURE LIMITING REAGENTS (PART 2)**

You still need to figure how many moles of MgCl2 you can make if you have 10.0 grams of magnesium and 20.0 grams of chloride!

1. Use dimensional analysis to solve:

20.0 grams Cl/35.45 g Cl/mol = .56 mols Cl.

.56 mols Cl/2 mols Cl • MgCL2) = .28 mols MgCL2

Since you have .56 moles of chloride and you need 2 moles of chloride to every 1 mole of magnesium you can make .28 moles of magnesium chloride.