# 

# Chatfield’s

# Chemistry Workbook

Fall 2019 Edition

Mr. Bertelsen and Mr. Fitch

This Book Belongs to:

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

# To the Student

This course is probably very different from any other class that you have taken. Your success will depend on your knowledge of science and math. You will find that this book is also different from any other that you have used.

You will be given the book in sections, and the book is yours to keep. You are encouraged to write in the book and highlight sections.

This book is not meant to take the place of the teacher. It does not contain all of the information that you need to answer the questions. You must attend class, take notes, participate in class discussions, and be active during lab work. This book attempts to summarize the information, to hit the high points of what you have already seen in class.

Our Thanks

To the students in our classes who helped us develop these materials and who pointed out numerous errors to us.

# Constants

Charge on an electron qe = 1.6 x 10-19 C

Mass of an electron me = 9.1 x 10-31 kg

Atomic Mass Unit amu = 1.66 x 10-27 kg

Avogadro’s Number n = 6.022 x 1023 mol-1

Universal Gas Constant R = 0.0821 (L atm) / (mol K)

Absolute Zero – 273 °C or 0 Kelvins

Planks Constant h = 6.63 x 10-34 J s

Speed of light c = 3 x 108 m/s

# Metric Summary

## Prefixes

pico = 10-12 = p nano = 10-9 = n micro = 10-6 = 

milli = 10-3 = m centi = 10-2 = c

kilo = 103 = k Mega = 106 = M Giga = 109 = G

## Common SI Units and conversions

Length: centimeters or meters conversion: 100 cm = 1 m

Mass: grams or kilograms conversion: 1000 g = 1 kg

Volume: Liters, milliliters, or cm3 conversion: 1 L = 1000 ml

1 ml = 1 cm3

Temperature: °C or Kelvins conversion: back of periodic table

# Exponents

The form 2.3E+5 means the same as 2.3 x 10+5 = 230,000

The form 1.6E-3 means the same as 1.6 x 10-3 = 0.0016

When multiplying, add exponents.

When dividing, subtract exponents.

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## What You Should Learn in Chemistry – Semester 1

**Unit 1 – The Basics… By the end of this unit, you should:**

1. Know the rules of scientific notation, significant figures, the metric system and unit conversions.
2. Know and understand the history of Chemistry including the progression from Thales to our current theories.
3. Understand the parts of the atom, where they are located and how they are used to determine chemical properties.
4. Understand chemical and physical changes and properties, including the differences between the 3 main states of matter – solid, liquid and gas.
5. Understand isotopes and relative abundance and be able to use this knowledge to determine the average atomic mass for an element.

**Unit 2 – The Periodic Table… By the end of this unit, you should:**

1. Understand electron orbitals, electron spin and the concept of valence electrons and their use in determining the electron configuration and electron energy level diagram for an element.
2. Identify groups (families) on the periodic table and understand the similarities within a group, trends within a group, and the reasons for these.
3. Identify periods on the periodic table and understand trends within each period and the reasons for these trends.
4. Understand the distinction between metals, metalloids and nonmetals and where they are found on the periodic table.
5. Be able to identify the alkali metals, alkaline earth metals, halogens, noble gases, transition metals, and the Lanthanide and Actinide series on the periodic table.
6. Understand the significance of valence electrons and be able to determine the number of valence electrons for an element or group.

**Unit 3 – Formulas and Equations… By the end of this unit, you should:**

1. Define the mole and be able to use Avogadro’s Number to convert between moles and particles of a substance.
2. Understand chemical bonding, including a distinction between ionic and covalent bonding and be able to determine the type of bond present in a compound based on the type of elements involved.
3. Be able to predict the electrical charge for ions based on their location on the periodic table.
4. Be able to name and write chemical formulas for ionic and molecular compounds.
5. Be able to write and balance chemical equations, including being able to predict products from given reactants, and to identify the type of reaction.
6. Be able to use the activity series to predict when a reaction will take place when given a set of reactants.

**Unit 4 – Stoichiometry… By the end of this unit, you should:**

1. Understand the significance of the mole within a balanced equation.
2. Be able to determine the molar mass of a compound and use this in conversions.
3. Understand the process of using a given amount of one substance to determine the amount of a different substance in a chemical equation.
4. Be able to do mass-to-mass problems, mass-to-mole problems, mole-to-mass problems, and mole-to-mole problems.
5. Understand limiting reactants and be able to determine the amount of product that will be produced and the amount of excess reactant that remains.

**Unit 5 – Energy in Chemical Reactions… By the end of this unit, you should:**

1. Understand enthalpy diagrams and be able to use them to identify if a reaction is endothermic or exothermic.
2. Understand the role of catalysts in regard to the activation energy of a given process.
3. Understand the change in energy associated with chemical processes.
4. Understand and be able to calculate the energy associated with changes of state.
5. Understand and be able to calculate the energy associated with the changes in temperature for a substance using the equation q = c m t

## Course Expectations

1. **Participation** is very important in this class, therefore attendance is vital. Beginning with your 4th unexcused tardy, you will lose 1% off your final grade for each tardy. Beginning with your 1st unexcused absence, you will lose 3% off your final grade. These can be made up after school: ½ hour for each tardy, and 1 hour for each absence

2. **Grading Percentages:**

Tests and Quizzes 50%

Final Exam 20%

Lab work 20%

Homework 10%

1. **Homework:** The purpose of homework is to allow you to practice the computations that we are learning and to reinforce key concepts that we cover. An important aspect of this is to ‘think in writing’. What this means is, while solving a problem, write out the steps that you take to solve it and then box your solution. All of your homework is to be done in a comp book, which I will check once each unit for your homework grade. We will discuss this more in class, but the 3 things that I will expect of all of your homework assignments.

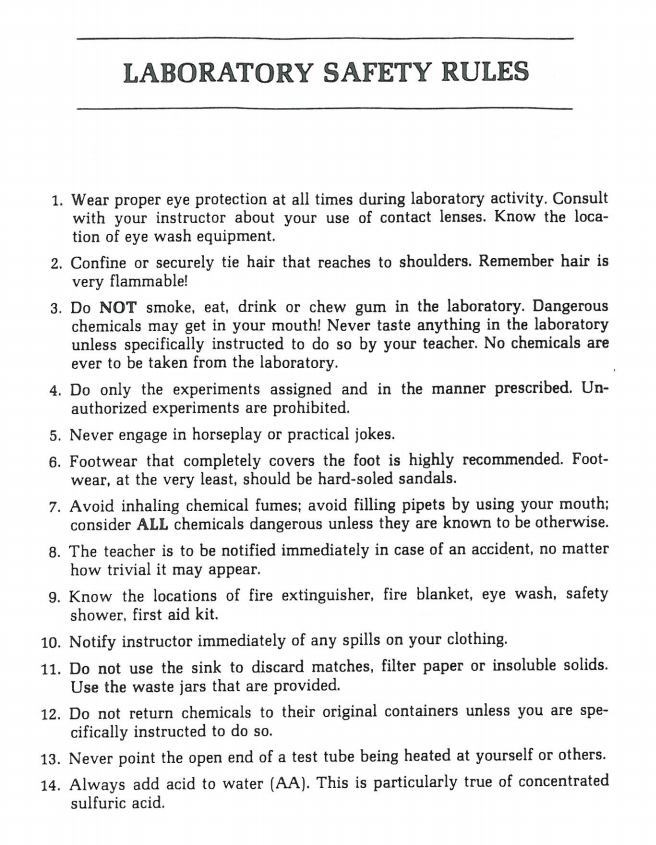
1. Do all work in your homework notebook (exceptions will be noted as we go).

2. Show all work that you do – if you use your calculator, write it down.

3. Box your final answer

1. **Labs:** All lab work must be made up within the first week after the lab has taken place. After this, the experiment will be put away and you will not have an opportunity to make up that lab. If you fail to make up the lab in the week, you will still be responsible for the material covered by the experiment.
2. **LATE HOMEWORK AND LABS ARE NOT ACCEPTED FOR FULL CREDIT**. After the due date for an assignment, you will receive 75% within the first week, and 50% after that. Papers turned in at the end of the semester to try to ‘catch up’ may be subject to an additional deduction.
3. **The Semester Grade** will be cumulative. This means that each term grade will simply be a progress report.
4. **THERE WILL BE NO MAKE-UP TESTS!!** Exceptions to this rule will be made on an individual basis. If a student is going to be absent on test day, arrangements must be made in advance to take the test early. To accommodate for the strictness of this policy, each person’s lowest test score will be dropped at the end of the semester.
5. **FINAL EXAM:** The final exam will count for 20% of your final grade, and will not be dropped even if it happens to be your lowest test grade. The final exam will be cumulative.
6. **Make Up Work:** If you have missed a day of class, it is your responsibility to determine what you missed and to make up the work. A calendar on my web site will assist in this. Along with the schedule will be a basket for returned papers and a file containing any assignment you may have missed. You may also find it helpful to get the notes for the day from a friend. You should only come to us after you have done the above mentioned things. *NOTE!!* The passing period is *NOT* the time to ask questions about things you may have missed. If you have questions, please come in during access for help with your questions.
7. **Everyone** will be responsible for monitoring their own learning. Only you know if you understand the material, and therefore you need to come in for help when aspects of chemistry are confusing to you. You are expected to meet the standard on lab assignments by receiving at least a 70%. If you do not accomplish this, you will have an opportunity to earn up to an 80%. In order to receive this chance, you will need to make an appointment to come in for help within two days of receiving the assignment back. If you miss your scheduled appointment, the penalty will be decided individually.
8. **Take Pride in Your Work!** Whenever you do an assignment, it will be important that you turn in your best effort. This means taking pride in everything that you do. Each assignment must include your full name, your class period, the date and the title of the assignment. When you turn in a sloppy paper, it tells me that you don’t care. This is okay, but you won’t get the same grade as someone who does in fact care. (A 30% reduction is possible)
9. **Class.** You will be expected to come to class prepared. The items that you will need include a three-ring notebook with blank paper and dividers, something with which to write and a calculator. During class you will be expected to take notes, and to use these notes as a reference tool. Many of the problems that we will learn early in the course will be problems repeated throughout, and you will need to be able to reference these types of problems.
10. **Showing Your Work.** You will be required to show your work on all problems that you do for this class. On homework, if you turn in an assignment without work (only answers) you will a zero for the assignment. In addition, your work needs to make sense. If I can’t follow what you are doing, I won’t be able to give you any partial credit.
11. **Odds and Ends…** This is a college prep course and many things will reflect that. You will have a greater level of freedom and a greater level of responsibility. Please be aware that it is your responsibility to come to class prepared. You will also be expected to come to class and get ready to take notes first thing. Once you have your notebook out and ready, then you may chat with friends until attendance is complete.
12. **Homework:** This is another area of increased responsibility. Homework is a learning tool. Getting the answer correct the first time is not the goal – understanding the problem is. Because of this, you will be grading your own homework during the course, and keeping each in a separate section of your notebook. On the day before each exam, a homework grade will be given based on the completeness of your work. If you have completed the problems with all of your work shown, you will receive full credit. Its correctness will not be assessed. Answers will be supplied during work days or can be found on line.
13. **Food.** With very few exceptions, you will not be able to bring food or drink to class, nor will you be allowed to leave class to buy these things
14. **Electronic Devices:** Cell phones are not allowed to be out or used in class except as approved or in an emergency and with teacher approval. Ipods or other musical device should only be used during work time. The general rule: your phone should not be out while trying to learn Chemistry. Any abuse of these guidelines may result in your device being confiscated for the remainder of the day. *On the day of a quiz or an exam, you are not allowed to have any electronic device out at any time during the entire class period. If you have such an item out on the day of a quiz or an exam, the device will be confiscated, and you will receive a 0% for that assessment.*
15. Safety: Everyone will be required to read and sign the safety contract that follow on the next two pages. It is provided here so that you will always have a copy for reference.





# Unit 1 – The History of Chemistry, Working with Numbers and The Properties of Atoms

## Algebra Review

1. Solve for the variable in each of the following equations:
   1. 300 = 273 + °C d. 24/n = 12
   2. 5x = 450 e. 16 = m/4
   3. m = 14 + (2 · 1) + (16 · 3) f. 3x + 4 = 10
2. Solve for the specified variable:
   1. PV = nRT (for P) d. K = 273 + °C (for °C)
   2. D = m/v (for m) e. °F = (9/5)°C + 32 (for °C)
   3. D = m/v (for v) f. Q = mcΔt (for c)
3. Given the equation and quantities, substitute and solve for the unknown quantities.
   1. D = m/v m = 32 g, v = 44.8 liters

Solve for D

* 1. P1V1 = P2V2 P1 = 1 atm, V1 = 10 liters, P2 = 2 atm

Solve for V2

* 1. PV = nRT P = 1 atm, V = 22.4 liters, R = 0.0821 L atm/mol K, T = 273 K

Solve for n

* 1. V1/T1 = V2/T2 V1 = 10 liters, V2 = 5 liters, T2 = 300 Kelvin

Solve for T1

* 1. Q = cmΔt c = 1 cal/g°C, m = 20 g, Δt = 20 °C

Solve for Q

* 1. Molar mass = m/n m = 24 g, molar mass = 12 g/mol

Solve for n

## Metric System

We will work exclusively in the metric system in this course, a measurement system with which few of us are completely comfortable. This review is designed to help you if you find that this system is difficult.

Within the metric system are two common systems of measurements. One you don’t need to know about for this course (it’s called the cgs system, which stands for centimeters, grams and seconds). The system that we will use is the **MKS** system. MKS stands for Meters, Kilograms and Seconds, the units we will use for distance, mass and time, respectively. Everything that measure during the course of this year will be in terms of these three quantities. If for some reason a measurement is taken using some other system, we will need to convert it into the MKS system. For this reason, understanding the basics of the metric system is quite important.

We start with the prefixed used in the metric system. These prefixes are used throughout, and are used with any base unit.

**Prefixes Equivalent Value Scientific Abbreviation**

**Notation**

Tera = 1,000,000,000,000 = 1012 T

Giga = 1,000,000,000 = 109 G

Mega = 1,000,000 = 106 M

Kilo = 1,000 = 103 k

Hecta = 100 = 102 h

Deka = 10 = 101 da

**Basic** = 1 = 100

Deci = 0.1 = 10-1 d

Centi = 0.01 = 10-2 c

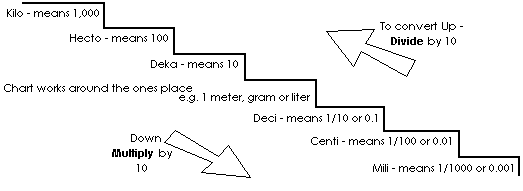
Milli = 0.001 = 10-3 m

Micro = 0.000 001 = 10-6 

Nano = 0.000 000 001 = 10-9 n

Pico = 0.000 000 000 001 = 10-12 p

When you need to convert within the metric system, (eg: meters to centimeters) you can use ratios or you can use the familiar metric stair-step. Moving to the left equates to dividing by 10 and moving to the right equates to multiplying by 10. If you move up the stairs, you move the decimal to the left and if you move down the stairs then you move the decimal to the right. So 1 km would be 1,000 m and 3 cm would be 0.03 m



## Are Significant Figures Important? A Fable

A student once needed a cube of metal which had to have a mass of 83 grams. He knew the density of this metal was 8.67 g/mL, which told him the cube's volume. Believing significant figures were invented just to make life difficult for chemistry students and had no practical use in the real world, he calculated the volume of the cube as 9.573 mL. He thus determined that the edge of the cube had to be 2.097 cm. He took his plans to the machine shop where his friend had the same type of work done the previous year. The shop foreman said, "Yes, we can make this according to your specifications - but it will be expensive."

"That's OK," replied the student. "It's important." He knew his friend has paid $35, and he had been given $50 out of the school's research budget to get the job done.

He returned the next day, expecting the job to be done. "Sorry," said the foreman. "We're still working on it. Try next week." Finally the day came, and our friend got his cube. It looked very, very smooth and shiny and beautiful in its velvet case. Seeing it, our hero had a premonition of disaster and became a bit nervous. But he summoned up enough courage to ask for the bill. "$500, and cheap at the price. We had a terrific job getting it right -- had to make three before we got one right."

"But--but--my friend paid only $35 for the same thing!"

"No. He wanted a cube 2.1 cm on an edge, and your specifications called for 2.097. We had yours roughed out to 2.1 that very afternoon, but it was the precision grinding and lapping to get it down to 2.097 which took so long and cost the big money. The first one we made was 2.089 on one edge when we got finished, so we had to scrap it. The second was closer, but still not what you specified. That's why the three tries."

"Oh!"

## Significant Figures – SigFigs

Every measurement we make has some inherent error due to the limitations of the measuring instrument and the experimenter. The numerical value recorded for a measurement should give some indication of the reliability (precision) of that measurement. After reading the rule, write the number of sig figs on the space provided. The answers are provided on the bottom of the page.

**Rules for Determining Significant Figures**

1. All nonzero digits are significant. 438 \_\_\_\_

* 1. \_\_\_\_

.653 \_\_\_\_

2. All zeros between two nonzero digits are significant 506 \_\_\_\_

* 1. \_\_\_\_

900.43 \_\_\_\_

3. Zeros to the right of a nonzero digit, but to the left of 4830 \_\_\_\_

an understood decimal point, are not significant. If 60 \_\_\_\_

such zeros are known to have been measured, they 4830. \_\_\_\_

are significant and should be specified as such by 60. \_\_\_\_

inserting a decimal point to the right of the zero 10200300 \_\_\_\_

4. In numbers less than one, zeros to the right of a 0.06 \_\_\_\_

decimal point that are to the left of the first non-zero 0.00407 \_\_\_\_

digit are never significant. They are place holders. 0.0056 \_\_\_\_

5. In numbers less than 1, the zero to the left of the 0.8 \_\_\_\_

decimal point is never significant. It is there to make 0.104 \_\_\_\_

sure that the decimal point is not overlooked. 0.002 \_\_\_\_

6. All zeros to the right of a decimal point and to the 8.0 \_\_\_\_

right of a nonzero digit are significant. 16.40 \_\_\_\_

* 1. \_\_\_\_

**Addition and Subtraction:**

Your answer can’t have more significant figures **after the decimal** than the smallest number of significant figures after the decimal in any of the numbers used to obtain the answer.

Examples: 12.5 cm + 0.135 = 12.635 cm. *In the correct number of sigfigs: 12.6 cm*

0.00354 s – 0.002 s = 0.00154 s = *0.002 s (correct number of sigfigs)*

**Multiplication and Division:**

Your answer can’t have more **total** significant figures than the smallest total number of significant figures in any of the numbers used in the calculation.

Examples: 2.5 m x 1.25 m = 3.125 m2 = *3.1 m2 (correct number of sigfigs)*

1000. m / 4.35 s = 229.8851 m/s = *230. m/s (correct number of sigfigs)*

7. a) 3.65 + 2.1 b) 4.25 – 3 c) 0.00023 + 0.0004

8. a) 3.60 x 12.74 b) 25.35 / 0.021 c) 175 x 3

Answers: 1) 3, 4, 3 2) 3, 7, 5 3) 3, 1, 4, 2, 6 4) 1, 3, 2 5) 1, 3, 1 6) 2, 4, 4 7) 5.8, 1, 0.0006 8) 45.9, 1200, 500

## Scientific Notation

In our work this year, we will sometimes work with numbers that are quite large, while at other times we will work with numbers that are very small. When we discuss the numbers of atoms in something like a glass of water, we will learn that there are a huge number of atoms! In a 12-ounce glass of water for example, there are over 11,800,000,000,000,000,000,000,000 water molecules, and since each molecule has 3 atoms, there are nearly 36,000,000,000,000,000,000,000,000 atoms in a glass of water! These numbers will be common for us to deal with and writing them will become a giant pain if we don’t come up with a short cut. Luck for us! We have a solution.

The way that we will deal with numbers like these is by a simple multiplication by 10. Let’s look at an example and see where it takes us. We start with the number 1.18 and multiply it by 10.

Now we do it again!!

Let’s keep going and see what happens...

and this can be written as:

If you think about it for a moment, you will see that each time we multiply our 1.18 by 10, we are in essence moving the decimal to the right one place, and this is the key to writing numbers in scientific notation. Let’s write the number of molecules in a glass of water again.

In order to get our little 1.18 to the number above, we would need to multiply by 10, 25 times! We could write that like this:

This is no savings at all, but using exponents, we can shorten this up as:

Each of the last 3 numbers that you see represent the same value, but the last one is the easiest to write, and that is why we use scientific notation. In order to take any number greater than 1 and write it in scientific notation, you begin by placing the decimal point between the first and second number. Once you have done this, count the number of decimal places to the right that you need to move the decimal, and this becomes the exponent on the 10. Let’s look at a couple of examples.

**Example1:** Write 135,000 meters in scientific notation.

We start by placing the decimal between the 1 and the 3 and get 1.35000

In order to get this number back to 135,000 we would need to move the decimal to the right 5 decimal places – which would get us back to the 135,000 – the result of this is that we can write the original number as:

Notice that we have the same number of sigfigs in this number as we had in the original number. Always make sure to keep the same number of sigfigs in the number written in scientific notation as you started with.

**Example 2:** Write 320.30 grams in scientific notation.

Again we start by placing the decimal between the first and second digit.

3.2030 – and then notice that we need to move the decimal 2 places to get back to the original number. The answer, with the same number of sigfigs, is:

In addition to these large numbers, we will also be using very small numbers and scientific notation can help us here too! Imagine one of those water molecules in our 12-ounce glass of water. Each water molecule has a mass of 18.02 atomic mass units – or amu for short. The amu is a unit that we will be using a fair amount this year. It is helpful when dealing on the atomic scale – that is, on the scale of atoms. Both the proton and the neutron have a mass of 1 amu. Let’s find out what this is in grams using a unit conversion and then we will write the number in scientific notation.

In order to covert from amu to grams we use the relationship between amu and grams. 1 atomic mass unit has a mass of 0.000000000000000000000001661 grams. To write this in scientific notation we again begin by placing the decimal between the first and second non zero digit:

Just as we multiplied by 10 to move the decimal to the right, we can divide by 10 in order to move the decimal to the left.

Dividing by 10 again would move the decimal one more space:

In order get from 1.661 to the value of 1 amu above, we need to divide by 10, 24 times. So again, we can write the value in three ways:

This last value, with the negative exponent comes from the notion that dividing by 10 is the same as multiplying by 10-1. Therefore, dividing by ten 24 times, is the same as multiplying by 10-24. Hopefully you can see that the last way is the easiest way to write this number. When you have numbers that are smaller than 1, you will use negative exponents for the powers of 10 in your numbers written in scientific notation. You still begin by placing the decimal between the first and second digit of your number, but instead of counting to the right you count to the left. In your answer the main difference will be that the exponent of the 10 will be negative. Perhaps a few examples will help.

**Example 3:** Write 0.00231 Liters in scientific notation.

We still begin by placing the decimal between the 2 and the 3:

In order to return to the original number we need to divide by 10 several times, specifically:

This number can be written more easily using a negative exponent for the division by 10

Returning to the amu example, we know that 1 amu = 1.661 x 10-24 g. We will use this now to determine the mass of our water molecule. Recall that 1 water molecule has a mass of 18.02 amu. To convert this to grams we set up the following conversion:

Our water molecule has a very small mass indeed! In order to type this into your calculator you will be using the EE button. This is a special button that takes the idea: ‘one point six six one times ten to the negative twenty fourth’ and makes it into one number. People new to the use of scientific notation often want to type the number in the conversion above (the number in parentheses) into their calculator as follows:

This is a mistake, as the calculator will treat this as a calculation, subject to the rules of order-of-operations and can cause problems - depending on the calculation. The correct way to type this number into your calculator is by using the EE button. When you do this, the calculator will show a capitol E in the display, which can be thought of as ‘times ten to the’. To properly type the conversion into your calculator using the EE button you should type:

When you do this, the calculator will display:

And the calculator will treat it as a number rather than a calculation. The EE button should be seen as your scientific notation friend – use it often!

**Homework Format** – The purpose of the assignments in this book is to allow you to practice the computations that we are learning and to reinforce key concepts that we cover. An important aspect of this is to ‘think in writing’. What this means is, while solving a problem, write out the steps that you take to solve it and then box your solution. One very important aspect of this is that you leave a trail of breadcrumbs – a path that you can follow backwards if, after checking your answer in the back of this book, you find that you have made a mistake. The homework assignments in this book are designed so that you can do all of your work in this book, but if you don’t show your work you will not get any credit. We will discuss this more in class, but the 3 things that I will expect of all of your homework assignments.

1. Do all work in your this workbook (exceptions will be noted as we go).

2. Show all work that you do – if you use your calculator, write it down.

3. Box your final answer.

4. FYI: there are 3 nautical miles in a league and 6076 feet in a nautical mile

## Homework #1 – Conversions, Sig Figs, and Scientific Notation

Part 1 – Perform the following conversions:

1. 3.85 m to cm
2. 35 ml to liters
3. 125 g to kg
4. 1.3 L to mL
5. 45 mL to cm3
6. 25.0 cal to kcal
7. 45 in to cm
8. 2.65 mL to fluid ounces
9. 20000 leagues to miles
10. 15 L to gallons
11. 45 Celsius to Kelvin
12. 15 ft to cm
13. 75 miles to km
14. 5280 ft to miles

Part 2 – How many significant figures do the following numbers have?

1. 68,945,510,000
2. 4,635,852,224
3. 0.009200
4. 201
5. 986
6. 0.0023405

Part 3 – Rewrite the following numbers using just two significant figures – you may need to write them in scientific notation to accomplish this.

1. 1,584
2. 50
3. 900,000,000,564
4. 987
5. 0.00206
6. 8.45

Part 4 – Write the following numbers in scientific notation. (be sure to keep the same number of sig figs)

1. 0.00234
2. 345.0
3. 760,908,001
4. 300,000,000 (speed of light m/s)
5. 0.0000000000000000000000000000009109 (mass of an electron in kg)
6. 602214000000000000000000 (number of atoms in one mole… we will learn about this later)

## Homework #2 – working with numbers and density

Part 1: Combining sig figs and scientific notation. For each of the following, determine the number of sig figs and write in scientific notation.

1. 4000 g
2. 0.0019 g
3. 50.05 g
4. 58900 ft
5. 58900.0 ft
6. 0.0120 km

Part 2: Round off each of the following to the indicated number of sig figs (indicated in parentheses) and write the number in scientific notation.

1. 2.68 g (2)
2. 47.374 ml (4)
3. 24 km (1)
4. 0.4851 in (2)
5. 0.06350 ml (2)
6. 0.002300 g (3)

Before beginning part 3, take a moment to read the rules for calculations using sigfigs.

**Addition and Subtraction:**

Your answer can’t have more significant figures **after the decimal** than the smallest number of significant figures after the decimal in any of the numbers used to obtain the answer.

Examples: 12.5 cm + 0.135 = 12.635 cm. *In the correct number of sigfigs: 12.6 cm*

0.00354 s – 0.002 s = 0.00154 s = *0.002 s (correct number of sigfigs)*

**Multiplication and Division:**

Your answer can’t have more **total** significant figures than the smallest total number of significant figures in any of the numbers used in the calculation.

Examples: 2.5 m x 1.25 m = 3.125 m2 = *3.1 m2 (correct number of sigfigs)*

1000. m / 4.35 s = 229.8851 m/s = *230. m/s (correct number of sigfigs)*

Part 3: Perform the following calculations using the appropriate number of sig figs:

1. 31.2 \* 580
2. 412000 \* 0.005
3. 82.250/0.0012
4. 8 / 51
5. 58 + 8.2
6. 100 + 78
7. 74.3 - 0.21
8. 6.02 x 1023 \* 8
9. 8.25 x 10-4 + 830
10. 4.28 x 105 / 5.0 x 104

Part 4 – Density Calculations.

1. Diamonds (a form of carbon) are measured in carats, and 1 carat = 0.200 g. What is the mass of a 5 carat diamond?
2. If the density of diamond is 3.51 g/cm3, what would the volume of the 5 carat diamond from the last problem?
3. The volume of a diamond is found to be 2.8 mL. What is the mass of the diamond in carats?
4. Hydrogen gas (H2) at 0C and 1 atmosphere pressure has a density of 0.0899 g/L. What is the density in g/cm3?
5. What volume would 2.02 grams of hydrogen occupy?
6. What mass of hydrogen would be required to fill the Hinderberg? (the volume of the great air ship was 2 x 105 cubic meters)
7. The density of aluminum is 2.70 g/cm3. Express this value in units of kilograms per cubic meter?
8. The diameter of a hydrogen nucleus is 1.0 X 10-15 meters and its mass is 1.67 X 10-24 grams. What is the density of the nucleus in g/cm3?

## Homework #3 – History and more…

Part I – History

1. Who was the first to propose that everything is composed of atoms?
2. What led Rutherford to his conclusions regarding the nucleus?
3. What is Bohr’s model of the atom and what led him to this conception?
4. What was Dalton’s atomic theory of matter?
5. What is the difference between qualitative and quantitative observations?
6. Why are different colored lines seen in the emission spectra of hydrogen?

Part II – In the space provided, mark each of the following as to whether it is a physical or chemical change.

1\_\_\_\_\_\_\_ A candle burning

2\_\_\_\_\_\_\_ Nitric oxide (a colorless gas) mixes with air to form a brown gas

3\_\_\_\_\_\_\_ Dry ice (CO2) sublimes (changes from a solid to a gas).

4\_\_\_\_\_\_\_ Two colorless liquids are mixed and a white precipitate forms

5\_\_\_\_\_\_\_ Steam condenses on a bathroom mirror

6\_\_\_\_\_\_\_ Gasoline explodes in a combustion engine

7\_\_\_\_\_\_\_ Healing of a wound

8\_\_\_\_\_\_\_ Digestion of food

9\_\_\_\_\_\_\_ Growth of a plant

10\_\_\_\_\_\_ Melting of ice

11\_\_\_\_\_\_ Formation of clouds in the air

12\_\_\_\_\_\_ Making of rock candy by evaporation of water from a sugar solution

13\_\_\_\_\_\_ Excavating of the earth

14\_\_\_\_\_\_ Explosion of TNT (trinitrotoluene)

15\_\_\_\_\_\_ Kicking of football

Part III – What are the five indicators of a chemical change?

1)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_2)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_3)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_4)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_5)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

True or False: If you have something you didn’t have before, a new compound or a new substance, chances are good you have a chemical change. Explain.

## Homework #4 – Parts of the atom… A Strain Your Brain!

For this assignment, you may use any resources available. Your brain, your neighbor (if they are willing), or your own book. Sorry no phones allowed! FYI – your teacher will not be very helpful on this… If we give you the answers – it isn’t much of a strain is it?...

1. The mass of an atom is often given in units of amu. What does this stand for?
2. What part (or parts) of the atom has a mass of 1 amu?
3. What is the mass of a proton, neutron and electron in amu?
4. What is the charge on an electron? A proton? A neutron?
5. What parts of the atom are found in the nucleus?
6. Where are the electrons found in the atom?
7. What does the atomic number of an element tell you?
8. What is the mass number of an atom?
9. What is the mass number of an atom with 6 protons and 6 neutrons?
10. Define or describe an ion.
11. Sodium has 11 protons. If there are 12 neutrons in a sodium atom, what is its atomic number?
12. What is the mass number of the atom in problem number 11?
13. Define or describe “isotope”.
14. Specifically, what subatomic particle causes the differences in mass between 235U (uranium 235) and 239U (uranium 239)?
15. If the mass number is always a whole number, why is the mass of potassium listed on your periodic table as 39.10 amu?
16. Look at your periodic table and copy the information for neon. Label the atomic number and the atomic mass on the following picture.

Ne

1. What is the atomic number for lithium? For copper? For beryllium?
2. What is the atomic mass of helium? Of zinc? Of aluminum?
3. What is the chemical symbol for sodium? For potassium? For magnesium?
4. What is an ion?
5. How many protons does fluorine have? How many does cobalt have? Boron?
6. If a fluorine atom has 10 electrons what is the charge on the atom?

## Homework #5 – More SigFigs, Density and Average Atomic Mass

*Here is some useful information ~ always think to check here, you will find goodies!*

*2.54 cm = 1 in, 2 cups = 1 pint, 1 mile = 5280 ft, 12 in = 1 ft*

Part 1: density and conversions revisited!

1. A glass container weighs 48.462 g. A sample of 4.00 ml of antifreeze is added and the container plus antifreeze weighs 54.513 g. Calculate the density of the antifreeze solution.
2. A 5.7 mg piece of gold is hammered into gold leaf of uniform thickness, with an area of 44.6 cm2. What is the thickness of the gold leaf? (the density of gold is 19.3 g/cm3)
3. A box with square base measuring .8 m per side has a height of 1.2 m. It is filled with 3.2 kg of expanded polystyrene packing material. What is the bulk density of the packing material, in g/cm3?
4. Hylon VII, a starch-based substitute for polystyrene packing material has a bulk density of 12.8 kg/m3. How many grams of the material are needed to fill a volume of 2.00 ft3?
5. Which of the following would most difficult to lift into the back of a pickup truck? (I) a 100 lb bag of potatoes (II) a 15 gallon plastic bottle filled with water or (III) a 3.0 L flask filled with mercury (density of water = 1g/ml; density of Hg = 13.534 g/ml)

Part 2: Parts of the atom:

…One thing to remember, if an atom has no charge then the number of protons will equal the number of electrons.

1. How many protons are there in carbon 12?
2. How many neutrons are there in sodium 23?
3. If an oxygen atom is neutral, how many electrons does it have?
4. How many protons are there in uranium 235?
5. How many neutrons are there in uranium 235?
6. What is the average atomic mass of Mercury?

Part 3: Average Atomic Mass:

1. A new element Cornium (Cn) has been discovered with two isotopes, Sweet Cornium and Feed Cornium. sweetCn has a mass of 0.82 amu and feedCn has a mass of 0.78 amu. If the relative abundance of sweet Cornium is 35.8% and the relative abundance of feed Cornium is 64.2%, determine the average atomic mass for Cornium.
2. Close on the heels of Corniums great discovery was the discovery of Beanium (Bn). Beanium has 3 isotopes, pintoBn, kidneyBn and BlackEyedBn. The masses of each are 1.25 amu, 1.36 amu and 0.98 amu respectively. If each has a relative abundance of 33.3% what is the average atomic mass for Beanium?
3. Carbon 12 occurs in nature 98.89% of the time, and carbon 13 occurs 1.11% of the time. What is the average atomic mass of carbon?

12C = 12.0000 amu 13C = 13.0034 amu

1. Lithium 6 has a relative abundance of 7.42% and lithium 7 a relative abundance of 92.58%. What is the average atomic mass of lithium?

6L = 6.0151 amu 7L = 7.0160 amu

1. The fractional abundance of nitrogen 14 is 99.63% and for nitrogen 15 is 0.37%. What is the average atomic mass?

14N = 14.0031 amu 15N = 15.0001 amu

## Unit 1 – Exam Review

Complete all activities and answer all questions to begin preparation for the Unit 1 Exam. You should also know your notes and the labs that we have done.

1. Understand the contributions of Dalton, Thompson, Rutherford and Bohr to our understanding of atomic structure, including the experiments that helped them to their conclusions.
2. What is the difference between qualitative and quantitative observations?
3. What are the five indicators of a chemical change?
4. Mark each of the following as to whether it is a physical or chemical change:
   1. \_\_\_\_\_\_ Dry ice (CO2) sublimes (changes from a solid to a gas).
   2. \_\_\_\_\_\_ Two colorless liquids are mixed and a white precipitate forms
   3. \_\_\_\_\_\_ Steam condenses on a bathroom mirror
   4. \_\_\_\_\_\_ A candle burning
   5. \_\_\_\_\_\_ Melting of ice
   6. \_\_\_\_\_\_ Formation of clouds in the air
5. What is the mass of a proton, neutron and electron in amu?
6. What is the charge on an electron? A proton? A neutron?
7. What parts of the atom are found in the nucleus?
8. Where are the electrons found in the atom?
9. What does the atomic number of an element tell you?
10. What is the mass number of an atom?
11. Define or describe an ion.
12. Define or describe “isotope”.
13. Specifically, what subatomic particle causes the differences in mass between 235U (uranium 235) and 239U (uranium 239)?
14. If atomic masses are always whole numbers, why is the mass of potassium listed on your periodic table as 39.10 amu?
15. What is the atomic number for iodine? For zinc? For calcium?
16. What is the atomic mass of neon? Of copper? Of tin?
17. What is the chemical symbol for gold? For tungsten? For silver?
18. How many protons does chlorine have? How many does cobalt have? Boron?
19. If a sodium atom has 10 electrons what is the charge on the atom?
20. How many significant figures in the following?
    1. 5000.05 g
    2. 65.42 g
    3. 5612 kg
    4. 0.047 cm3
    5. 58,900,000 ft
21. Round off the following to the indicated sig figs.
    1. .0165 L to 3 sig figs
    2. 24 km to 1 sig fig
    3. 0.4851 to 2 sig figs
    4. 268 g to 2 sig figs
    5. 47.374 ml to 2 sig figs
    6. 8.34987 ml to 3 sig figs
22. Perform the following calculations using the appropriate number of sig figs:
    1. 9.47 \* 220
    2. 812.000 \* 0.0050
    3. 82.250 / 12
    4. 18 / 515
    5. 58 + 8.2
    6. 0.05 + .100
    7. 100 + 78
    8. 74.3 - 0.2113
23. Perform the following conversions:
    1. 3.85 cm to m
    2. 7.5 L to mL
    3. 90.25 g to kg
    4. 13 mL to L
    5. 45 mL to cm3
    6. 22 °C to Kelvin
    7. 32 ft to cm
    8. 30 miles to Km
    9. 160 Km to miles
24. A 25 lb. anchor is made of pure iron. It displaces 1.62 quarts of water when used. What is the density of the iron in g/cm3?
25. What is the density of a 4.085 g steel ball that has a 5.5 mm radius? (volume of a sphere is 4/3πr3)
26. The density of liquid water at 22 ºC is 1g/ml. The density of solid water is .917 g/ml. What is the volume of 145 ml of liquid water after it freezes? What was the gain in volume due to solidification?
27. What mass of methanol (density = .791 g/ml) occupies the same volume as 15.0 kg of gasoline (density = .690 g/ml)?
28. A box with square base measuring .8 m per side has a height of 1.2 m. It is filled with 3.2 kg of expanded polystyrene packing material. What is the bulk density of the packing material, in g/cm3?
29. Which of the following would most difficult to lift into the back of a pickup truck? (I) a 100 lb bag of potatoes (II) a 15 gallon plastic bottle filled with water or (III) a 3.0 L flask filled with mercury (density of water = 1g/ml; density of Hg = 13.534 g/ml)
30. How many protons are there in nitrogen 14 (14N)?
31. How many neutrons are there in potassium 40 (40K)?
32. If a calcium atom is neutral, how many electrons does it have?
33. How many protons are there in plutonium 244?
34. How many neutrons are there in plutonium 244?
35. What is the average atomic mass of tungsten? Tin? Tellurium?
36. Mr. Cox has been doing some research and has discovered a new element – Gravitonium (Gv). Using a Ouija board, he discovered that there are two isotopes, Strong and Weak Gv, with masses strongGv = 298.7 amu and weakGv = 15.86 amu. If the strong isotope accounts for 180 out of every 200 Gv atoms, what is the average atomic mass for Gravitonium?
37. Mrs. Valerio was envious of Mr. Cox’s work and decided to find her own element. Her element, Mathmatium (Mt) has an average atomic mass of 89.5 amu.
38. Carbon 12 occurs in nature 98.89% of the time, and carbon 13 occurs 1.11% of the time. What is the average atomic mass of carbon?

12C = 12.0000 amu 13C = 13.0034 amu

1. Lithium 6 has a relative abundance of 7.42% and lithium 7 a relative abundance of 92.58%. What is the average atomic mass of lithium?

6L = 6.0151 amu 7L = 7.0160 amu

1. The fractional abundance of nitrogen 14 is 99.63% and for nitrogen 15 is 0.37%. What is the average atomic mass?

14N = 14.0031 amu 15N = 15.0001 amu

1. Chlorine 35 has a relative abundance of 75.53% and chlorine 37, 24.47%. What is the average atomic mass?

35Cl = 34.9689 amu 37Cl = 36.9659 amu

1. Uranium 235 occurs 0.72% of the time and uranium 238 occurs the other 99.27%. What is the average atomic mass?

235U = 235.0439 amu 238U = 238.0508 amu

# Unit 2 – The Periodic Table

## Homework #1: Ain’t nuttin’ but trends ~ A Strain Your Brain

Your mission: Using the data table below and your periodic table, find 5 trends that exist within the periodic table. Once you have found them, your group will need to visually represent one of these trends. You may choose a graph, a chart or other means, but it must be visual and the trend must be clear.

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Element symbol | Atomic number | Atomic mass | Boiling point (K) | Melting point (K) | Valence electrons | Oxide | Chloride |
| H | 1 | 1.01 | 20.268 | 14.025 | 1 | H2O | HCl |
| He | 2 | 4.00 | 4.215 | 0.95 | 2 | –– | –– |
| Li | 3 | 6.94 | 1615 | 453.7 | 1 | Li2O | LiCl |
| Be | 4 | 9.01 | 2745 | 1560 | 2 | BeO | BeCl2 |
| B | 5 | 10.81 | 4275 | 2300 | 3 | B2O3 | BCl3 |
| C | 6 | 12.01 | 4470 | 4100 | 4 | CO2 | CCl4 |
| N | 7 | 14.01 | 77.35 | 63.14 | 5 | N2O5 | NCl3 |
| O | 8 | 15.99 | 90.18 | 50.35 | 6 | –– | OCl2 |
| F | 9 | 18.99 | 84.95 | 53.48 | 7 | OF2 | FCl |
| Ne | 10 | 20.18 | 27.096 | 24.553 | 8 | –– | –– |
| Na | 11 | 22.99 | 1156 | 371.0 | 1 | Na2O | NaCl |
| Mg | 12 | 24.31 | 1363 | 922 | 2 | MgO | MgCl2 |
| Al | 13 | 26.98 | 2793 | 933.25 | 3 | Al2O3 | AlCl3 |
| Si | 14 | 28.09 | 3540 | 1685 | 4 | SiO2 | SiCl4 |
| P | 15 | 30.97 | 550 | 317.30 | 5 | P2O5 | PCl3 |
| S | 16 | 32.06 | 717.75 | 388.36 | 6 | SO3 | SCl2 |
| Cl | 17 | 35.45 | 239.1 | 172.16 | 7 | OCl2 | –– |
| Ar | 18 | 39.95 | 87.30 | 83.81 | 8 | –– | –– |
| K | 19 | 39.10 | 1032 | 336.35 | 1 | K2O | KCl |
| Ca | 20 | 40.08 | 1757 | 1112 | 2 | CaO | CaCl2 |

Once you have found five trends, answer the following questions.

1. What are your five trends?
2. Show your visual representation of one of the trends. This should include a graph of some sort. *Please note – you will not receive credit for an atomic number vs. atomic mass plot.*
3. What would the formula be for Barium oxide? Lead Chloride?

The first group finished with the questions may present their work to the class for extra points. If you decide to do this, you will need to give a few-minute presentation outlining the trends that you found, and the thought process that brought you to finding them.

## Homework #2 – And Now…More Trends! ~ A Strain Your Brain

Please read through this assignment before you begin. You will find that there are some statements. Do not be troubled – these are true and should be read and understood to help with previous and subsequent questions. Chapter 5 in Modern Chemistry may be very helpful in your quest for understanding! Please do not use your phone… only your book!

**Part I:** Some more trends.

1. What happens to the size of atoms as you go down a group? Why
2. Electronegativity is a measure of how strongly an atom can attract an electron to itself. The higher the electronegativity, the more it wants to gain an electron.
3. What happens to the size of the atoms as you move across a period? Why?
4. The attraction between the electron and the proton decreases as they get farther apart.
5. Does the size of an atom change when electrons are added? Why or why not?
6. Does the size of an atom change when electrons are lost? Why or why not?
7. What are valence electrons? How many valence electrons are there in the alkali metals? The noble gases? The halogens? The alkaline earth metals?
8. Adding or removing electrons changes the size of an atom because they take up space.
9. What is the general trend in ion size as you move across a period? Down a family? Why does this happen?

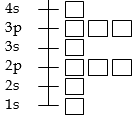
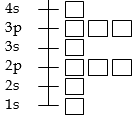
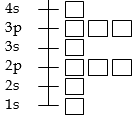
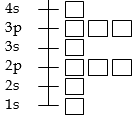
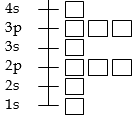
**Part II:** And now for some more trends…

1. What is meant by an atom’s ionization energy? What is the trend in ionization energies as you move across a period? Why?
2. What is the trend in ionization energy as you move down a family? Why?
3. What is meant by an atom’s electronegativity? What is the trend in electronegativity as you move across a period?
4. New orbitals take up space, but as the atomic number within a period grows, the number of protons grows and this exerts more force on the electrons.
5. What is the trend in electronegativity as you move down a family?
6. Why are the halogens likely to form ions with a -1 charge?
7. When an atom in the alkali metals ionizes (gains or loses e-) what will it’s new charge be?
8. What is the goal of all elements in the kingdom of ert?

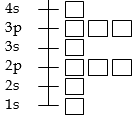
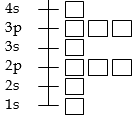
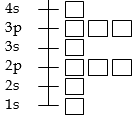
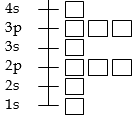
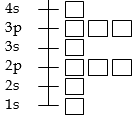
## Homework #3 – Electron configuration and energy level diagrams

Begin by writing the electron configuration for the element listed. Once you have the configuration, fill in the energy level diagram with the appropriate arrows to indicate the electrons in that energy level and suborbital.

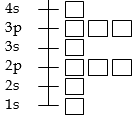
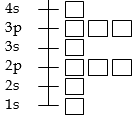
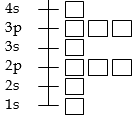
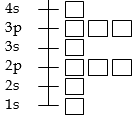
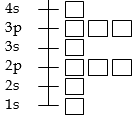
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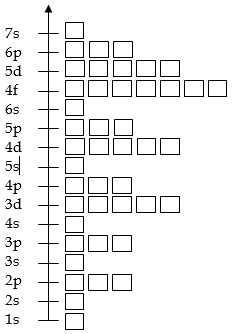
N O F Ne Na



Al Si P Cl Ca



Given the complete diagram, finish by drawing your own diagrams and filling them in.



Ti As Rn Ba Hg

## Homework #4 – More energy level diagrams

Begin again by writing the electron configuration for the ions listed. Once you have the configuration, draw an energy level diagram with the appropriate boxes and draw the diagram for the ion. If it helps, you can just draw lines for your energy levels instead of boxes. Make sure that your arrows are in the correct direction for each suborbital.

N3- Mg2+ Cl- Al3+

K+ Ti4+ Fe3+ Ge4+

Kr I- Ag+ Ba2+

# I Get A’s On My Homework – But I still Bomb the Tests!...

Some students feel they do all the homework but that they don’t do well on the exams. Most of these students have done well in school, but school really hasn’t been very hard – do the work, look over the notes, take the test, earn a good grade. For many students this strategy no longer works, and frustration sets in. Learning how to overcome this frustration and how to study for tests may be the most important skills you get from this class.

Chemistry is a branch of science, and science uses mathematical concepts to explain the relationships of matter. It is imperative to develop the skills necessary to manipulate the algebraic expressions that are used so often in this course.

I think most students need to work problems to prepare for chemistry exams. I believe they need to **work the problems before looking at notes or examples.** Since exams don’t come with examples, you need to be able to work problems without having a pattern to follow. Here are some recommendations:

* Look at the problems in the homework packet that the exam will cover. Notice when there are changes in directions or level of difficulty. For each change, pick at least one problem to work. Pick problems for which you have answers. Remember that about 70% of the test will be C level questions – basic problems, 10% - 15% will be B level problem – integrating a couple concepts, and 10% - 15% will be A level questions – integrating and extending concepts that we have covered.
* Work some example patterns without the answers by your side. Understand what problems challenged you the most, and figure out why. Once you have finished a couple of problems, check your answers and identify which ones you missed. Now is the time to look at notes and examples. Find some time before the exam to go back and try these problems again. Most students who have tried the above find that they did not know as much math as they thought they did. Working problems like this usually identifies what math you own and what math you don’t own.

**Radical Concept**

You don’t have to wait until the night before a test to study. If you have a quiz on Wednesday and do most of your studying on Monday, you have Tuesday’s class to get help and Tuesday night to rework any problems you missed on Monday.

Word problems are hard to study. As you do a series of word problems, pick the two or three that you think are the most typical of the assignment. Since we usually have labs that demonstrate or mimic real life examples, use your lab write ups to review these real-world problems as well. Too often students complete the labs to get a result - they don’t consider that the labs are designed to demonstrate the process of problem solving. If you have trouble understanding the labs, then chances are any word problem will be a challenge as well. If the exam includes word problems (and it usually does), go back and set up various examples and problems similar to the ones asked in labs.

Tests should be easier to prepare for than quizzes because I’ve already given you samples of what I think are the most important concepts. **Make sure you look at (rework?) any quizzes over material that a test will cover.**

**For most students, success in chemistry results from hard work, fighting through the frustration, and:**

* **Completing assignments on time**
* **Getting questions answered as soon as possible**
* **Good preparation for exams**
* **Paying attention to details (especially arithmetic)**

## Unit 2 – Exam review

Part 1: General Vocabulary – define the following:

Valence electrons

Period

Family or Group

Halogens

Noble gases

Alkali metals

Alkaline Earth metals

Transition metals

Metals, nonmetals and metalloids

Lanthanide/Actinide series

Electron spin

s, p, d and f block

Part 2: Draw the energy level diagram, write the electron configuration, and state the number of valence electrons for each of the following:

N F-1 Mg+2 O-2 Al+3 Se-1

Part 3: Name the element that has the electron configuration:

1. [Ar]4s2 3d7
2. [Rn] 7s2 5f3
3. [Kr] 5s2 with a charge of +2
4. [He] 2s2 2p6 with a charge of –2
5. 1s2 2s2 2p6 3s2 3p6 4s2 3d10 4p6 5s2 4d10 5p2
6. 1s0 with a charge +1

Part 4: Sketch the s, px, py, pz orbitals.

Part 5: Answer the following questions.

1. How many d orbitals are there?
2. How many f orbitals are there?
3. One day while walking near the beach, a sodium atom lost an electron to a chlorine atom. What is the new electron configuration for the chlorine and the sodium atom?
4. Write the electron configuration for each of the following
5. The smallest halogen atom
6. The alkali metal with a full 3p orbital
7. The three lightest alkaline earth metals
8. The element similar in chemical behavior to Ga in the same period as Sn
9. The nonmetallic element in the same family lead
10. The (as yet undiscovered) noble gas after radon
11. Why are the group 1 metals (the alkali metals) so reactive?
12. Which of the halogens is most reactive? Why?
13. Which alkaline earth metal is most reactive? Why?
14. What orbital is being filled in the alkaline earth metals?
15. Why does the 4s orbital fill up before the 3d orbital?
16. Write the order in which the electrons go into the orbitals from 1s to 7p, including the number of electrons which can fit into each orbital.
17. What elements are classified as metalloids? Why are they classified as this?
18. How does the atomic radius change as you move down a group? As you move across a period?
19. How does the size of an atom change when electrons are added? Why?
20. How does the size of an atom change when electrons are lost? Why?
21. What is the definition of the ionization energy for an atom?
22. How does the ionization energy change as you move across the periodic table?
23. What is a mole of atoms?
24. What is meant by the electronegativity of an atom?
25. Which atom has the greatest electronegativity? The smallest?
26. Why can there be only 2 electrons in any given orbital?

Extra Credit: How many potassium atoms will react with one oxygen atom?

Draw the 3dxy orbital. Draw the electron energy level diagram for the as yet undiscovered noble gas after radon.

# Unit 3 – Writing Formulas and Chemical Equations

## Homework #1 – Naming Chemical Compounds ~ a strain your brain

Your Mission: the following list shows several chemical compounds and their names. From this list, formulate a list of rules that would enable you to name additional compounds. (…the back of your periodic table may help with some of the names)

NaCl Sodium Chloride

NaOH Sodium Hydroxide

CaO Calcium Oxide

MgF2 Magnesium Fluoride

Zn3N2 Zinc Nitride

Zn(NO3)2 Zinc Nitrate

BaSO4 Barium Sulfate

SrS Strontium Sulfide

Al2Te3 Aluminum Telluride

NH4OH Ammonium Hydroxide

Li3PO4 Lithium Phosphate

K4C Potassium Carbide

H2CO3 Hydrogen Carbonate

Na3P Sodium Phosphide

1. What do all of the above compounds have in common?
2. What are your rules for naming the compounds?
3. What would the name be of the following compounds?

Ra(OH)2 NH4Cl MgBr2

## Determining the Molecular Formula of Compounds

There are two general types of compounds that we have looked at.

***Ionic Compounds***

You will know you have an ionic compound if you have a metal and a nonmetal, or if the compound has one or more polyatomic ion.

1) Determine the charge on each ion.

* Metals – charge is determined by the element's location on the periodic table.

Alkali Metals: 1+

Alkaline Earth Metals: 2+

Transition Metals: Roman numerals tell you the charge or you can look on the back of your periodic table for the ion

* Non-metals – charge is determined by the element's location on the periodic table.

Halogens: nearly always 1-

Oxygen’s Family: usually 2-

Nitrogen’s Family: often 3-

Polyatomic Ions: Charge found on the back of your periodic table

2) Balance the charge by adjusting the number of each ion. You want the total amount of positive and negative charge to be equal.

Subscripts are used to designate the number of each ion in the compound.

Example: Chromium (III) Carbonate

step1: Cr3+ determined from the Roman numeral.

CO32- found on the back of the periodic table

step 2: The charges will balance with a total positive of 6+ and a total negative of 6-.

To get this charge we will need 2 chromium and 3 carbonate ions.

***Answer: Cr2(CO3)3***

***Molecular Compounds***

If you have 2 or more non=metals in the compound, then you have a molecular compound.

The number of each element in the compound will be designated by the use of prefixes, and the elements will be written in the order that they appear. The prefixes can be found on the front of your periodic table. If the first element in the compound is singular, the ‘mono’ prefix is usually omitted. For example, CO2 is usually just called carbon dioxide rather than monocarbon dioxide, while CO is only know as carbon monoxide.

Example: Nitrogen triiodide – one nitrogen and 3 iodine.

***Answer: NI3***

## Homework #2 – Naming Chemical Compounds

1) Write the name of each of the following Binary compounds:

MgS\_\_\_\_\_\_\_\_\_\_\_ KF\_\_\_\_\_\_\_\_\_\_\_ BaI2\_\_\_\_\_\_\_\_\_\_\_ Li2O\_\_\_\_\_\_\_\_\_\_\_\_ ZnO\_\_\_\_\_\_\_\_\_\_\_ VO2\_\_\_\_\_\_\_\_\_\_ KCl\_\_\_\_\_\_\_\_\_\_\_\_ BeBr2\_\_\_\_\_\_\_\_\_\_

SrO\_\_\_\_\_\_\_\_\_\_\_\_ Mg2C\_\_\_\_\_\_\_\_\_ CaF\_\_\_\_\_\_\_\_\_\_\_ Rb2S\_\_\_\_\_\_\_\_\_\_\_

1. Write the names for the following ternary compounds. (ternary compounds are formed by three different elements)

K2Cr2O7\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Ba(OH)2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Al(NO3)3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Li2CO3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Zn3(PO4)2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CaSO4\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ V(OH)2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Pb2O3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Zn(OH)2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ AgNO3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Zr(CrO4)2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Cs3PO4\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Binary Ionic Compounds: Write the chemical formulas for the following compounds.

Barium Fluoride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Aluminum Sulfide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Lithium Oxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Potassium Chloride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Magnesium Bromide \_\_\_\_\_\_\_\_\_\_\_ Zinc (II) Chloride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Sodium Iodide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Radium Phosphide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Cesium Nitride\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Calcium Oxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Hydrogen sometimes forms bonds with various nonmetals, making an acid. We will be using the following 6 acids enough that you will need to memorize their common names:

Hydrochloric Acid HCl Nitric Acid HNO3

Sulfuric Acid H2SO4 Phosphoric Acid H3PO4

Acetic Acid (Vinegar) HC2H3O2 Carbonic Acid H2CO3

1. Some of the metals can form more than one type of ion. This is true for transition metals and for metals in the lanthanide and actinide series. For example: Iron can form an ion with a +2 or a +3 charge, and in the ion name, roman numerals are used to designate the charge. In this case, Fe+2 is written Iron (II) whereas Fe+3 is written Iron (III). Write the formulas for the following compounds.

Copper (II) chloride \_\_\_\_\_\_\_\_\_\_\_\_\_ Chromium (III) Oxide \_\_\_\_\_\_\_\_\_\_\_\_\_

Lead (II) Bromide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Nickel (IV) Sulfide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Vanadium (IV) Iodide \_\_\_\_\_\_\_\_\_\_\_\_ Iron (III) Nitride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Plutonium (IV) Chloride\_\_\_\_\_\_\_\_\_\_ Lead (II) Phosphide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Molybdenum (VI) Fluoride\_\_\_\_\_\_\_\_ Uranium (VI) Oxide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Ternary compounds are compounds comprised of three different elements. Write the empirical formula (its simplest form) for the following ternary compounds.

Zinc (II) Sulfate \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Copper (II) Carbonate \_\_\_\_\_\_\_\_\_\_\_\_\_\_

Barium Chromate \_\_\_\_\_\_\_\_\_\_\_\_\_\_ Potassium Phosphate \_\_\_\_\_\_\_\_\_\_\_\_\_\_

Aluminum Hydroxide \_\_\_\_\_\_\_\_\_\_\_ Calcium Oxalate \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Sodium dichromate \_\_\_\_\_\_\_\_\_\_\_\_ Lead (IV) Nitrate\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Iron (III) Nitrate \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Ammonium Sulfate \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Write the correct formula or name for the following compounds:

Sulfur dioxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ NaHCO3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Lanthanum (III) perchlorate\_\_\_\_\_\_\_\_\_\_\_\_ CaCr2O7­\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Sulfur trioxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ HCl\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Carbon dioxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ PCl3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Hydrogen Iodide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ OF2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Carbon tetrachloride \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ HBr\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Dihydrogen monoxide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ TiO2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Potassium permanganate\_\_\_\_\_\_\_\_\_\_\_\_\_\_ NaC2H3O2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Nitric Acid \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CrN\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Plutonium (IV) oxide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ Ar\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Nitrogen disulfide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ H2SO4\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Vinegar \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ H2O2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Carbon monoxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ NCl3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Acetic Acid\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ HNO3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Phosphoric Acid\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CdTe\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

Potassium thiocyanate\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ V(NO2)5\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. There are several polyatomic ions that occur often enough that memorizing them is helpful. Please memorize the following ions:

Ammonium NH4+ Nitrate NO3-

Hydroxide OH- Carbonate CO32-

Chromate CrO42- Sulfate SO42-

Dichromate Cr2O72- Phosphate PO43-

Acetate C2H3O2-

1. When two or more nonmetals form a compound, we call it a molecular compound. In these compounds, the prefixes mono, di, tri, tetra, penta, hexa, hepta, octa, nona and deca are used to designate the number of each element in the compound. For example, SO2 is sulfur dioxide. Using the correct prefixes, name the following:

CO\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ NO2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ SO3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ CCl4\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ PCl5\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ NH3 (ammonia)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

CH4 (methane)\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ SO2\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

NI­3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ SF3\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

1. Molecular formulas (for covalent compounds) are easy to write when given the name. Simply use the number word roots to tell you how much of each element is there. If no number word root is listed, assume there’s one (mono isn’t always used).

antimony tribromide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ hexaboron silicide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

chlorine dioxide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ hydrogen iodide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

iodine pentafluoride\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ dinitrogen trioxide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

ammonia \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ phosphorus triiodide\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

oxygen\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ selenium hexafluoride\_\_\_\_\_\_\_\_\_\_\_\_\_\_

disilicon hexafluoride\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ sulfur tetrachloride\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

methane\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ diboron silicide \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_

nitrogen trifluoride\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ tetraphosphorus pentasulfide\_\_\_\_\_\_\_

## Homework #3 - More naming…gettin’ ready for the quiz

Part 1:

Name the following compounds.

1. MgO
2. NI3
3. NaCl
4. Fr3N
5. CO2
6. KCl
7. Ca3(PO4)2
8. CH4
9. NO2
10. AlCl3
11. Zn(NO3)2
12. AgOH
13. CsC2H3O2
14. Cu2O
15. PbBr4
16. HgO
17. He
18. ZrO2

Part 2:

Write the molecular formula for the following compounds

1. Nitrogen dioxide
2. Carbon trifluoride
3. Magnesium nitrate
4. Cobalt (II) phosphide
5. Molybdenum (VI) carbonate
6. Beryllium bromate
7. sodium chloride
8. Ammonium
9. Sodium bicarbonate
10. Radium oxide
11. Potassium permanganate
12. Manganese (IV) oxide
13. Dicarbon tetrafluoride (teflon)
14. Titanium (III) sulfite
15. Uranium (VI) oxide
16. Zinc (II) perchlorate
17. Aluminum borate
18. Chromium (III) acetate

Part 3: Answer the following general questions:

1. How many valence electrons do the following elements have?
2. Boron
3. Carbon
4. Nitrogen
5. Oxygen
6. Fluorine
7. Tellurium
8. Barium
9. Sodium
10. Germanium
11. Write the symbol and the typical charge if the element becomes an ion.
12. Sulfur
13. Tellurium
14. Rubidium
15. Nickel (II)
16. Niobium (V)
17. Plutonium (IV)
18. Bismuth (III)
19. Astatine
20. Selelium
21. What happens to an element when it ionizes?
22. What orbitals are being filled in the transition metals?
23. What orbitals are being filled in the Lanthanide and Actinide series?
24. Which elements are metalloids?
25. Write the formula for the eight polyatomic ions you are to memorize.
26. What is the octet rule?

Part 4: Draw the Lewis Dot Diagrams for the following:

1. F
2. H
3. He
4. CO2
5. Al
6. Ne
7. N2
8. O2
9. Sb-3
10. Ammonia
11. CCl4
12. Methane
13. PO43-
14. CO32-
15. SO42-
16. Ammonium
17. Nitrogen triiodide
18. Teflon

## The Big naming Quiz review

Name the following compounds:

1. PO4-
2. NO2

3. H2SO4

4. FeCl3

1. HCl
2. Li2SO4
3. PF5
4. Cu(NO3)2
5. KSCN

Write the formulas for the following compounds:

1. barium hydroxide
2. ammonium
3. lead (IV) carbonate
4. silicon tetrafluoride
5. phosphoric acid
6. sodium phosphate
7. aluminum sulfide
8. acetate
9. nitric acid
10. Write the Lewis Dot Diagram for Cl and Cl-
11. Write the Lewis Dot Diagram for carbon tetrachloride
12. How can you tell if a compound is ionic compound? How can you tell if a compound is a covalent (molecular) compound?
13. When metals ionize, they tend to form what kind of ions?
14. Describe a polyatomic ion.
15. How many valence electrons do the alkali metals have? The alkaline earth metals? The oxygen family? Halogens?
16. Describe the octet rule.

## Writing Equations and Predicting Products

As we begin working to write chemical equations, a summary of the process may be helpful. This outline should be used as needed while working on the following homework sets.

1. Determine the formula for each compound in the reaction
2. Once the formula for each compound is determined, reactants are written on the left side of the equation and products are written on the right side.
3. If the products are not given, you must decide what products will form. At this level, the compounds in the products will only be ionic compounds.

Some things to keep in mind when predicting products:

* Anions will not bond with anions. (ex: NO31- will not bond with OH1-)
* Cations will not bond with cations. (ex: Mg2+ will not bond with Na1+)
* Subscripts on the compounds do not necessarily stay the same.

ex: Na + Cl2 🡪 NaCl. The subscript of 2 on chlorine in the reactants does not stay 2 in the products because sodium chloride is written to balance the charge in this ionic compound

* Polyatomic ions will not break apart in a reaction unless you are told that they will.

The cation from one reactant will join with the anion of the other reactant, and visa – versa. These products must be written using the normal rules for writing compounds.

Example: Magnesium hydroxide reacts with phosphoric acid.

1) Magnesium hydroxide: Mg(OH)2

Phosphoric acid: H3PO4 – this is also known as hydrogen phosphate

1. Mg(OH)2 + H3PO4
2. Predicting products: From the activity series on the back of your periodic table (this will be covered in homework #6) we see that magnesium will react with acids, replacing the hydrogen in the hydrogen phosphate. Mg will bond with phosphate and hydrogen will bond with hydroxide. ie: Mg3(PO4)2 and HOH or H2O

Mg(OH)2 + H3PO4 🡪 Mg3(PO4)2 + H2O

4) The reaction should next be balanced.

3 Mg(OH)2 + 2 H3PO4 🡪 Mg3(PO4)2 + 6 H2O

5) The process is now complete.

## Homework #4 – Equation Writing

Part 1: Translate the chemical equation from the molecular names to the chemical symbols for each compound.

1. iron + sulfur 🡪 iron (II) sulfide
2. magnesium + oxygen 🡪 magnesium oxide
3. chlorine + lithium bromide 🡪 lithium chloride + bromine
4. sodium + chlorine 🡪 sodium chloride
5. copper (II) sulfate + iron 🡪 iron (II) sulfate + copper
6. hydrochloric acid + zinc 🡪 zinc chloride + hydrogen gas
7. sulfuric acid + iron 🡪 hydrogen gas + iron (II) sulfate
8. sodium chloride + potassium nitrate 🡪 sodium nitrate + potassium chloride
9. aluminum + oxygen 🡪 aluminum oxide
10. sulfuric acid + barium chloride 🡪 barium sulfate + hydrochloric acid
11. hydrogen + oxygen 🡪 water
12. sodium phosphate + calcium carbonate 🡪 sodium carbonate + calcium phosphate

Now go back and label each equation as one of the four reaction types. (Single Replacement, Double Replacement, Synthesis or Decomposition)

Part Two – Translate the word equations into equations with the chemical formulas

Also include the type of the reaction as you did at the end of part one.

1. Iron (III) oxide reacting with hydrochloric acid to produce iron (III) chloride and water.
2. Calcium hydroxide reacts with carbon dioxide to produce calcium carbonate plus water.
3. Aluminum sulfate plus sodium hydroxide produces aluminum hydroxide and sodium sulfate.
4. Zinc hydroxide reacting with sulfuric acid to produce zinc sulfate and water.
5. Iron metal and chlorine gas reacting to give iron (III) chloride.
6. Sodium hydroxide plus sulfuric acid will react to produce water plus sodium sulfate.
7. Bismuth (III) sulfide and oxygen gas produces bismuth (III) oxide and sulfur dioxide gas.
8. The element barium reacting with water to give barium hydroxide and hydrogen gas.
9. Diatomic bromine reacting with sodium iodide to produce sodium bromide and diatomic iodine molecules.
10. A molecule of propane (C3H8) reacting with oxygen gas to give carbon dioxide gas and water.

## Homework #5 – Equation Writing and Balancing

Part One – Translate the chemical equation from the molecular names to the chemical symbols for each compound. Once you have written the equation, please balance it.

1. iron + nitrogen 🡪 iron (II) nitride
2. sodium + oxygen 🡪 sodium oxide
3. ammonia 🡪 hydrogen + nitrogen
4. copper (II) carbonate + iron 🡪 iron (III) carbonate + copper
5. sulfuric acid + nickel 🡪 hydrogen gas + nickel (III) sulfate

Now go back and label each equation as one of the four reaction types. (Single

Replacement, Double Replacement, Synthesis or Decomposition)

**Part Two** – Translate the word equations into equations with the chemical formulas

Also include the type of the reaction as you did at the end of part one. Once you have the equation written, please balance it.

1. Iron (III) sulfide reacting with hydrochloric acid to produce iron (III) chloride and dihydrogen sulfide.
2. Calcium hydroxide reacts with carbon dioxide to produce calcium carbonate plus water.
3. Aluminum sulfite plus sodium hydroxide produces aluminum hydroxide and sodium sulfite.
4. Zinc hydroxide reacting with nitric acid to produce zinc nitrate and water.
5. Copper metal and chlorine gas reacting to give copper (I) chloride.

## Homework #6 – Predicting Products in Reactions

Part One: predicting products. Given the reactants, predict what the products will be and then write the equation using the correct formulas for each compound. The first five give the type of the reaction. For 6 – 12, please specify the reaction type.

1. DR nitric acid + potassium hydroxide 🡪
2. Decomposition mercury (II) oxide 🡪
3. Synthesis sodium + chloride 🡪
4. SR aluminum + hydrochloric acid 🡪
5. Decomposition sodium chloride decomposing into sodium and chlorine gas
6. barium hydroxide + sulfuric acid 🡪
7. copper (II) chloride + lead 🡪 (assume lead (IV) in the products)
8. sodium iodide + bromine 🡪
9. nitrogen + hydrogen 🡪
10. hydrogen + oxygen 🡪
11. sulfuric acid + calcium hydroxide 🡪
12. strontium phosphate + sodium oxide 🡪

Part Two: Writing and balancing equations. Write the complete balanced equation for each reaction, including the reaction type.

1. Ammonia, NH3, decomposing into its component elements (both products exist as diatomic molecules)
2. Tungsten carbide, WC, and oxygen gas reacting to give tungsten (VI) oxide and carbon dioxide gas.
3. Ammonium dichromate decomposing to produce nitrogen gas, water and chromium (III) oxide.
4. Sulfur tetrafluoride reacting with water to yield sulfur dioxide and hydrogen fluoride.
5. Ammonia and fluorine gas reacting to produce dinitrogen tetrafluoride and hydrogen gas.
6. Ethane, C2H4, burning in the oxygen of the air to give carbon dioxide and water.
7. Vanadium (V) oxide and hydrogen gas produces vanadium (III) oxide and water
8. Boron (III) oxide and carbon reacting to produce boron carbide, B4C and carbon monoxide
9. Aluminum sulfide and water produces aluminum hydroxide and dihydrogen sulfide.
10. Manganese (IV) oxide and hydrochloric acid reacting to give manganese (II) chloride, chlorine gas, and water.
11. Calcium and oxygen reacting in a synthesis reaction

Extra Credit: gasoline (C6H14) burning. (hint: consider the O2 as O O in a Double Replacement. Lewis dot diagrams can help. The carbon will form double bonds in this case.)

## Homework #7 – The Activity Series

Using the activity series, determine if the reaction will take place. If you determine that a reaction will take place, write the complete balanced equation.

1. Fe + CuNO3 🡪 (assume iron (II) in products)
2. Al + Ca(OH)2 🡪
3. Cl2 + NaBr 🡪
4. Mg + AgNO3
5. Li + HCl 🡪

For questions 6 – 15, first write the formula for the reactants, then determine if the reaction will take place. If it does, write the complete balanced equation.

1. Calcium is dropped into a container of boiling water
2. Nickel is mixed with silver chloride (use nickel (II))
3. Mercury comes into contact with a solution of magnesium nitrate
4. Zinc is mixed with a sodium chloride solution
5. Fluorine gas surrounds a sample of sodium chloride
6. Silver is placed into a beaker with sulfuric acid
7. Ammonium bromide mixes with iodine
8. Copper (II) nitrate slowly mixes with lithium
9. Water is poured onto a large piece of iron
10. Phosphoric acid and potassium are mixed

# Unit 4 – Stoichiometry

## Homework #1 – Conversion problems

I) Conversions are at the heart of Chemistry and so we need to practice a bunch! Please answer the following questions.

1. What is the atomic mass of zinc?
2. Methylene chloride (CH2Cl2) is used as a solvent in paint strippers. What is the formula mass of CH2Cl2?
3. What is the atomic mass of nitrogen?
4. Ammonia (NH3) is a common household cleaning agent. What is the formula mass of NH3?
5. Sodium Hypochlorite (NaClO) is the active ingredient I household bleach. what is the formula mass of NaClO?
6. What is the atomic mass of neon?
7. What is the formula mass of sodium chloride?
8. What is the atomic mass of uranium?
9. Nitric acid is a strong acid. What is the formula mass of nitric acid?
10. What is the atomic mass of silver?

II) Calculate the molar mass (the mass of 1 mole) for each of the following compounds.

1. sulfuric acid
2. potassium nitrate
3. aluminum chloride
4. copper (II) sulfide
5. mercury (II) dichromate
6. carbon dioxide
7. carbon tetrachloride
8. methylamine (CH3NH2)
9. benzene (C6H6)
10. propane (C3H8)

III) Determine the mass in grams of each of the following

1. 1.00 mole of Li
2. 1.00 mole of Al
3. 1.5 mole of Ca
4. 6.02 X 1023 atoms of C
5. 12.04 X 1023 atoms of Ag
6. 3.0 mole Hg
7. 4.25 mole Cu
8. 1.37 mole Nickel (II) nitrate
9. 2.57 X 108 mole Sulfur dioxide
10. 1.75 mole Vanadium (V) oxide

IV) Determine the number of moles in each case.

1. 6.02 X 1023 atoms Ne
2. 2.25 X 1025 atoms Zn
3. 50 molecules barium hydroxide
4. 5.87 X 1023 atoms Ni
5. 4 X 1023 atoms Ca
6. 11.5 g Na
7. 2.65 g Fe
8. 0.0075 g O2
9. 3.25 X 105 g lead (II) hydroxide
10. 4.5 X 10-12 g Oxygen

V) Determine the mass, in grams, of the following

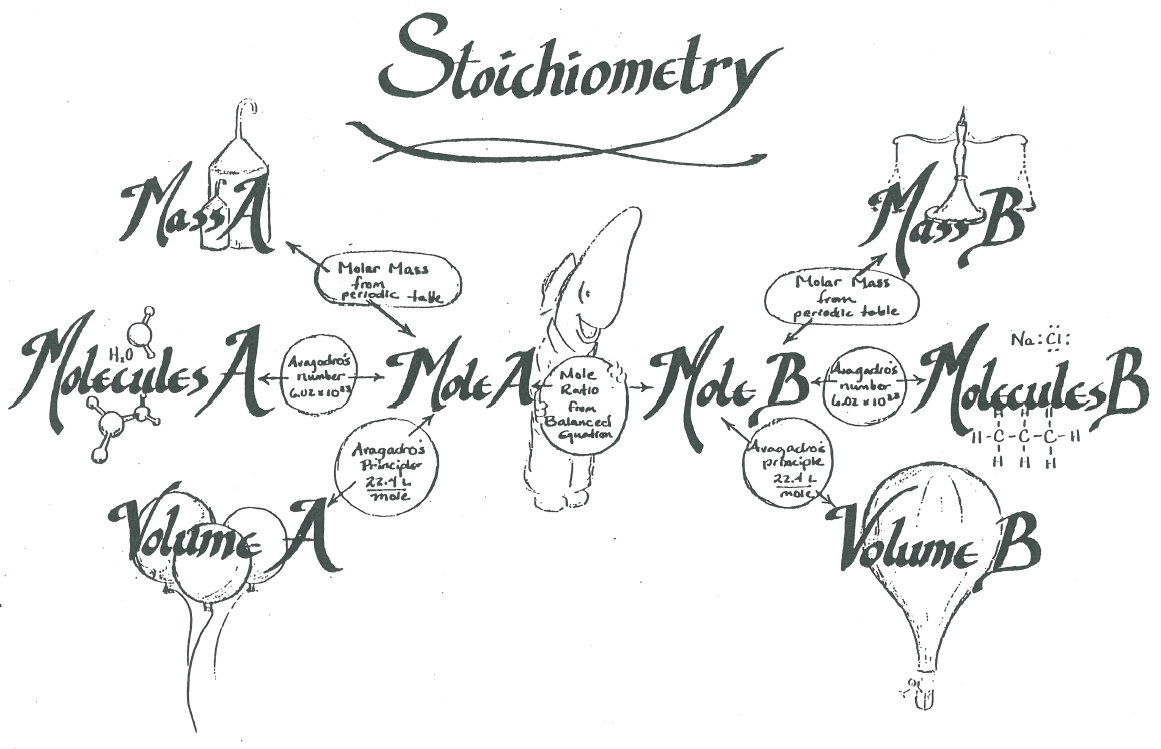
1. 0.5 mole F2
2. 1.50 X 1023 atoms Mg
3. 4.50 X 1012 atoms Cl
4. 8.42 X 1018 molecules Barium nitrate
5. 25 molecules tungsten (VI) oxide

VI) Determine the mass of the following

1. 2.5 mole sodium hydroxide
2. 1.5 mole nitric acid
3. 2.75 mole sulfur dioxide
4. 20 mole Lead
5. 5.4 X 1023 B atoms (hint: convert to moles first)
6. 8.02 X 1022 atoms S
7. 1.5 mol K
8. 0.02550 mol Pt
9. 1.00 X 10-10 mol Au
10. 5.0 kg Ammonium phosphate

VII) Review Problems:

1. How many atoms are there in 20 grams Lead (II) nitrate?
2. How many grams are in 5 mole of ammonia?
3. How many moles of hydrogen are there in 1 mole of water?
4. What is the mass of 3 moles of Sulfuric acid?
5. How many moles are there in 15 grams of zinc?
6. How many grams are there in 2 X 1023 atoms of mercury?
7. How many atoms are there is 350 grams of nitrogen trioxide?
8. How many moles are there in 180 grams of water?
9. How many moles of sulfur are there in 30 g of aluminum sulfate?
10. How many moles of aluminum?
11. How many grams of hydrogen are there in 4.5 moles of hydrogen gas?
12. If I gave you a 2-liter bottle filled with water, how many grams of water would you have?
13. How many moles of hydrogen is that?
14. How many atoms of hydrogen?



## Homework #2 – Stoichiometry problems

1. Consider the reaction between sodium and chlorine: 2Na + Cl2 🡪 2NaCl. 5 moles of sodium will produce how many moles of sodium chloride?
2. Water is produced in the following reaction: 2H2 + O2 🡪 2H2O
3. How many moles of hydrogen are needed to react with 10 moles of oxygen?
4. If you wanted 15 moles of water, how much oxygen should you start with?
5. What is the mass of the 15 moles of water produced in the previous problem?
6. For the reaction between iron and copper (II) sulfate: Fe + CuSO4 🡪 Cu + FeSO4, how many moles of copper will be formed from 15 moles of iron?
7. Given the chemical equation, how many moles of each reactant are needed to produce 2 moles NaOH?

Na2CO3 + Ca(OH)2 🡪 2NaOH + CaCO3

1. Ethane (C2H6) burns according to the *unbalanced* reaction

C2H6 + O2 🡪 CO2 + H2O

1. How many moles of oxygen are required for 4.5 moles of ethane?
2. How many moles of each product are formed?
3. Sulfuric acid reacts with sodium hydroxide according to the following:

H2SO4 + NaOH 🡪 Na2SO4 + H2O

1. Balance the equation for the reaction.
2. How many moles of sulfuric acid would be required to react with 0.75 mol NaOH?
3. What is the mass of this amount of sulfuric acid?
4. What mass of each product is formed in this reaction?
5. Sodium chloride is produced from its elements through a synthesis reaction.
6. Write and balance the equation for this process.
7. What mass of each reactant would be required to produce 25.0 mol of sodium chloride?
8. Copper reacts with silver nitrate through a single replacement reaction. (assume copper (II) will be the form of copper in the products) Write the balanced equation for this process.
9. If 2.25 grams of silver are produced from the reaction, how many moles of copper (II) nitrate are also produced?
10. How many moles of each reactant are required for the reaction?
11. Iron is produced from iron ore through the following unbalanced reaction in a blast furnace: Fe2O3(s) + CO(g) 🡪 Fe(s) + CO2(g) (s) = solid (g) = gas
12. If 4.0 kg of iron (III) oxide are available to react, how many moles of CO are needed?
13. How many moles of each product are formed?
14. Methanol (CH3OH) is an important industrial compound that is produced from the following unbalanced reaction: CO(g) + H2(g) 🡪 CH3OH(g). What mass of each reactant would be needed to produce 100.0 kg of methanol?
15. Nitrogen combines with oxygen in the atmosphere during lightning flashes to form nitrogen monoxide, which then reacts further with O2 to produce nitrogen dioxide. Write the two equations for this process.
16. What mass of NO2 is formed when NO reacts with 384 g O2?

1. How many grams of NO are required to react with this amount of O2?
2. As early as 1938, the use of NaOH was suggested as a means of removing CO2 from the cabin atmosphere of spacecraft according to the following unbalanced reaction: NaOH + CO2 🡪 Na2CO3 + H2O
3. If the average human body discharges 925.0 g of CO2 per day, how many moles of NaOH would be needed each day?
4. How many moles of each product would be formed each day?
5. The double-displacement reaction between silver nitrate and sodium bromide produces silver bromide, a component of photographic films.
6. If 4.5 mol of silver nitrate react, what mass of sodium bromide is required?
7. What mass of silver bromide is formed?
8. In a soda-acid fire extinguisher, concentrated sulfuric acid reacts with sodium bicarbonate to produce carbon dioxide, sodium sulfate and water.
9. How many moles of sodium bicarbonate would be needed to react with 150 g of sulfuric acid?
10. What mass of each product would be formed?

Extra Credit: Aspirin (C9H8O4) is produced through the following reaction of salicylic acid (C7H6O3) and acetic anhydride (C4H6O3): C7H6O3(s) + C4H6O3(l) 🡪 C9H8O4(s) + H2O(l). a) What mass of aspirin (in kg) could be produced from 75.0 mol of salicylic acid? b) What mass of acetic anhydride (in kg) would be required? c) Given the density of water, (1g/ml) how many liters of water would be formed?

## Homework #3 – Limiting Reactant Problems

1. Given the reactant amounts specified in each chemical equation, determine the limiting reactant in each case:

a) HCl + NaOH 🡪 NaCl + H2O

2.0 mol 2.5 mol

b) Zn + 2HCl 🡪 ZnCl2 + H2

2.5 mol 4.0 mol

c) 2Fe(OH)3 + 3H2SO4 🡪 Fe2(SO4)3 + 6H2O

4.0 mol 6.5 mol

1. For each reaction in the previous problem, determine the number of moles of excess reactant that remains.
2. For each reaction above, calculate the number of moles of each product formed
3. If 2.5 mol of copper and 5.5 mol of silver nitrate are available to reacts in a single displacement reaction, identify the limiting reactant. Assume Copper (II) in products.
4. Write the equation for this reaction.
5. Determine the number of moles of excess reactant remaining.
6. Determine the number of moles of each product formed.
7. Determine the mass of each product formed.
8. Sulfuric acid reacts with aluminum hydroxide in a double displacement reaction.
9. Write the equation for this reaction.
10. If 30.0 g of sulfuric acid react with 25.0 g of aluminum hydroxide, identify the limiting reactant.
11. Determine the mass of excess reactant remaining.
12. Determine the mass of each product formed.
13. A solution with 3.0 mol of silver nitrate is mixed with a solution containing 5.0 mol of sodium chloride.
14. Write the equation for this reaction.
15. Which reactant is the limiting reactant?
16. Which reactant is in excess?
17. How many moles of excess reactant are left?
18. How many moles of each product are produced?
19. In a container, 12.0 moles of nitrogen react with 12.0 moles of hydrogen producing ammonia (NH3).
20. Write the equation for this reaction.
21. Which reactant is the limiting reactant?
22. Which reactant is in excess?
23. How many moles of excess reactant are left?
24. How many moles of each product are produced?
25. One hot day, a container holding 10.0 moles of propane bursts into flames generating a cloud of carbon dioxide and water vapor.
26. Write the equation for this reaction.
27. Is there a limiting reactant in this case?
28. How many moles of each product are produced?
29. Calcium chloride reacts with sodium phosphate in a double displacement reaction.
30. Write the equation for this reaction.
31. If 100.0 g of calcium chloride is mixed with a solution containing 100.0 g of sodium phosphate, which reactant is the limiting reactant?
32. Which reactant is in excess?

1. How many moles of excess reactant are left?
2. How many grams of each product are produced?

As you have been working to understand the process of Stoichiometry, some of you may still be struggling. If you are finding it hard to work with unit fractions to plan your problems, a stoichiometry flow chart may be helpful. The following is provided to give you another way to conceptualize these problems and perhaps give you a structure that can help you in your quest to conquer the great Stoichiometry Mountain!

## Stoichiometry Outline

Reactant + Reactant 🡪 Product + Product

\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_

(grams reactant) (grams reactant) (grams product) (grams product)

\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_ \_\_\_\_\_\_\_\_\_\_

(moles reactant) (moles reactant) (moles product) (moles product)

Mole to Mole Conversions: Look at the **COEFFICIENTS ONLY!!!!** Place the coefficient of the substance you are going to on **TOP**. Place the coefficient of the substance you are coming from on the **BOTTOM**. You can only do this with a **BALANCED EQUATION**…*So BALANCE FIRST!!!!!*

## Stoichiometry Outline Example

To help understand the outline further an example may help. Suppose 28.5 g of aluminum hydroxide reacted with hydrochloric acid. How much water would be produced?

1 Al(OH)3 + 3 HCl 🡪 1 AlCl3 + 3 H2O

28.5 g . . 19.8 g .

0.365 mol Al(OH)3 3 mol H2O 1.10 mol H2O

1 mol Al(OH)3

If there is no coefficient on a given substance, then it is considered to be 1. You can make conversions going anywhere you want!!!! Notice how the units cancel as you go from moles Al(OH)3 to moles H2O. Use units! It helps.

# Unit 5 – Energy in Chemical Reactions

## Homework #1 – Temperature Changes and Heats of Reaction

1. Calculate the number of calories needed to raise the temperature of 200 grams of water from 10°C to 80°C.
2. Calculate the amount of heat needed to change the temperature of 200 grams of aluminum form 20°C to 70°C. (cAl = 0.22 cal / g°C)
3. A certain metal needed 75.0 calories of heat energy to raise the temperature of 10.0 grams of the metal from 25°C to 55°C. Calculate the specific heat of the metal.
4. A 40.00-gram block of iron is heated from 26°C to 160°C. How much heat does the iron block absorb? (cFe = 0.11 cal / g°C)
5. A 4.00-kg block of iron is heated from 26°C to 180°C. How much heat does it absorb? (see #4)

1. A 160-g glass cup at 20°C is plunged into hot dishwater at 80°C. If the cup’s temperature is raised to that of the dishwater, how much heat does the cup absorb? (cglass = 0.16 cal / g°C)
2. A beaker containing 120 mL of water is heated from 20°C to 70°C. How much heat does the water absorb?
3. How much heat is lost when a 4110-gram block of aluminum cools from 660°C to 25°C? (see #2)
4. How much heat is required to raise the temperature of 4.66 grams of carbon tetrachloride from 20.0°C to 77.0°C? (c = .201 cal/g°C)
5. The temperature of 250mL of ethanol (cethanol = 0.58 cal/g°C density = 0.8g/mL) increases from 10.2°C to 18.7°C in a calorimeter cup. How much heat did the ethanol absorb?
6. The cooling system of a car contains 20 liters of water. If the initial temperature is 20°C, what will be the final temperature is the engine operates until 200 kilocalories of heat are absorbed?
7. Suppose the cooling system from problem 5 is drained and filled with methanol at 20°C. (cmethanol = 0.59 cal / g°C density = 0.80 g/ml). What would be the final temperature of the methanol if it were to absorb 200 kilocalories of heat? (hint: use the density to determine mass)
8. Write a balanced equation for the melting of one mole of ice. Include the heat term. (H = 1.44 kcal)
9. Write a balanced equation for the melting of ice. Use the H notation. Is this an exothermic or endothermic process?
10. Write the balanced equation for the condensation of one mole of gaseous water. Include the heat term. (H = 9.7 kcal)
11. Write the balanced equation for the condensation of one mole of gaseous water. Use theH notation. Is this an exothermic or endothermic process?
12. When ethanol (C2H5OH) burns, 330 kcal/mol of energy is released. The two products that are formed are carbon dioxide and water. Write the balanced equation for this process, including the heat term. Is this an exothermic or endothermic process?

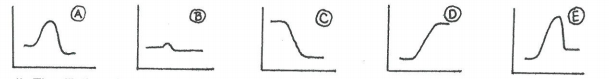
## Homework #2 – Heat Transfer and Change of State

1. How much heat is needed to change 50g of ice at O°C to water at O°C?

1. How much heat is needed to heat 50g of water at 0°C to water at 100°C?
2. How much heat is needed to change 50g of water at 100°C to steam at 100°C?
3. Assuming that the specific heat of steam is the same as water, how much energy is needed to heat 50g of steam from 100°C to 125°C?
4. Draw a heating curve for the temperature and phase changes outlined in problems 1-4. Label the graph with the appropriate melting and boiling temperatures
5. Consider 75 grams of solid sodium at its freezing point. How much heat would have to be absorbed by the sodium to melt, warm up to the boiling point and then completely vaporize? (c = 0.32 cal/g°C Hv = 1048 cal/g Hf = 27.4 cal/g Melting point = 98°C Boiling point = 889°C) careful… this is a three step problem!
6. Draw the heating curve for the heating process in problem 6. Label the graph with the appropriate melting and boiling temperatures.

## Homework #3 – Enthalpy Diagrams and Reaction Rates

Questions 1-4 refer to the enthalpy diagrams shown. In each instance, choose the diagram which best fits the described process.



1. The dilution of sulfuric acid by water is spontaneous, rapid reaction that is accompanied by the evolution of large amounts of heat.
2. Ammonia, NH3 may be made by the combination of nitrogen and hydrogen according to the following equation:

N2 + 3H2 🡪 2NH3 + 22 kcal

The use of catalysts, temperature of 5000C and pressure of 250 atm are necessary to get a yield of about 30% ammonia.

1. Carbon (graphite) may be converted to carbon (diamond) under extremely high temperature and high pressure. H = 37.5 cal/g.
2. Ammonium chloride dissolves readily and spontaneously in water. When this process occurs, the reaction vessel becomes cold.

Match the statements with the correct term(s).

5. The minimum energy necessary to start a chemical A. Activation energy

reaction B. Activated complex

6. A substance added to a chemical reaction to alter C. Catalyst

the reaction rate but is not consumed in the over D. Combustion

all reaction. E. Condensation

7. A biological catalyst F. Endothermic

8. A system that releases heat G. Enthalpy

9. Process of liquid changing to a gas H. Enzyme

10. Process of solid changing to a gas I. Evaporation

11. The total energy contained within a molecule J. Exothermic

12. A system that absorbs heat K. Initiation step

13. An intermediate molecule which forms before the L. Melting

final product and only lasts for a short period of time M. Sublimation

1. What is the latent heat of fusion for water?
2. What is the latent heat of vaporization for water?
3. When a substance melts, the temperature remains constant, even though you are adding heat to it. Where does the heat go?
4. Define exothermic and endothermic.
5. For each of the following processes, state whether they are endothermic or exothermic.
6. Wood burning in a fire place
7. Solidification of wax
8. H2O(l) + 9.7 kcal 🡪 H2O(g)
9. A piece of iron changing temperature from 80°C to 20°C
10. Water vapor condensing in the atmosphere (rain)
11. Burning gasoline in your car
12. H2O(l) 🡪 H­2O(s)+ 1.44 kcal

Draw the enthalpy diagrams for the following processes.

1. Ice melts absorbing 1.44 kcal/mol of heat.
2. Ethanol burns and produces 330 kcal/mol.
3. Water vaporizes absorbing 9.7 kcal/mol.
4. A 20-gram block of copper is heated from 100C to 750C. (hint: first calculate the heat necessary for this process…see HW 5-1)

Extra Credit: 15 grams of ethanol at 250C is heated to its boiling point (78.30C) and then vaporized (HV = 204.3 cal/g). Draw the enthalpy diagram for this process.

## Homework #4 – Energy in Reactions

1. Glucose is converted into energy in cellular respiration according to the following *unbalanced* equation. If 0.05 moles of glucose are converted to energy, how much heat is released?

C6H12O6 + O2 🡪 CO2 + H2O + 2043 kJ

1. Hydrogen peroxide (H2O2) decomposes into water and oxygen (H = -197 kJ). If 51 grams of hydrogen peroxide react, how much energy is released?
2. Aluminum can be produced from aluminum oxide according to the following equation: 2 Al2O3(s) + 3352 kJ 🡪 4 Al(s) + 3O2

How much aluminum oxide would be needed to produce 3 kg of aluminum?

How much energy would be needed?

1. Carbon dioxide is formed from carbon and oxygen by:

C(s) + O2(g) 🡪 CO2(g) H = -394 kJ/mol

How much energy is released as 35 grams of carbon are converted?

1. The heat of reaction for the combustion of 1 mol of ethyl alcohol (C2H5OH) is -9.50 x 102 kJ. How much heat is produced when 11.5 g of alcohol are burned?
2. The H for the complete combustion of 1 mol of propane is -2.2 x 103 kJ. Calculate the heat of reaction for the combustion of 33.0 g of propane (C3H8)?
3. The H for the formation of sulfur dioxide is -297 kJ/mol. How many kilojoules of energy are given off when 25.0 g of SO2 are produce from its elements?
4. The formation of ice from liquid water is an exothermic process given by the following equation: H2O (l) 🡪 H2O (s) + 6.03 kJ How much heat is released when 3.0 liters of water freeze?

Extra Credit: Your barbecue grill uses propane and can hold about 20 pounds of propane when the tank is full. If the energy required to melt ice is equal to the amount of energy released as ice is frozen, how many grams of ice could you melt with a full tank of propane?

## Some Final Thoughts and Equations to Help You Get Bye…

**Displacement** (units: meters) – how far an object travels

**Velocity** (units: m/s) – the rate of change of displacement. Sometimes called speed, but speed

is the scalar equivalent of velocity. Speed is the magnitude of velocity.

## Some of the equations we have used…

***Velocity equations:***  v = d = v t t =

***Acceleration equations:*** a = a =

# Answers

Measurement #1

1a. 4 1b. 3 1c. 2 1d. 4

1e. 3 1f. 3 1g. 5 2a. 13.6 cm

2b. 137 m 2c. −0.039 s 2d. 910 kg m/s2 2e. 2000 m/s

2f. 0.063 Nm 3a. 1.56 m 3b. 2.3 mm 3c. 1.2 kg

3d. 0.125 km 3e. 0.00250 ms 3f. 350. mg 4. 42

Unit 3 – HW#3 part 1

1. Magnesium oxide 2. Nitrogen triiodide 3. Sodium chloride 4. Difrancium nitride

5. Carbon dioxide 6. Potassium chloride 7. Calcium phosphate

8. Carbon tetrahydride 9. Nitrogen dioxide 10. Aluminum chloride 11. Zinc (II) nitrate

12. Silver (I) hydroxide 13. Cesium acetate 14. Copper (I) oxide 15. Lead (IV) bromide,

16. Mercury (II) oxide 17. Helium 18. Zirconium (IV) oxide

Unit 3 – HW#3 part 2

1.NO2 2. CF3 3. Mg(NO3)2 4. Co3P2

5. Mo(CO3)3 6. Be(BrO3)2 7. NaCl 8. NH4+1

9. NaHCO3 10. RaO 11. KMnO4 12. MnO2

13. C2F4 14. Ti2(SO3)3 15. UO3 16. Zn(ClO4)2

17. AlBO3 18. Cr(C2H3O2)3

Unit 3 – HW#3 part 3

1a. 3 1b. 4 1c. 5 1d. 6

1e. 7 1f. 6 1g. 2 1h. 1

1i. 4 2a. S2- 2b. Te2- 2c. Rb1+

2d. Ni2+ 2e. Nb5+ 2f. Pu4+ 2g. Bi3+

2h. At1- 2i. Se2- 3.Gains or loses e- 4. d orbitals

5. f orbitals 6. B, Si, Ge, As, Sb, Te, At

7. NO31-, CO32-, SO42-, PO43-, OH-, CrO42-, Cr2O72-, NH41+

8. atoms tend to gain, lose or share e- to complete their octet

Unit 3 – HW#4

1.kdk 2. fh

Unit 3 – HW#4

1.kdk 2. fh

Unit 3 – HW#4

1.kdk 2. fh

Unit 3 – HW#4

1.kdk 2. fh

Unit 3 – HW#4

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Unit 3 – HW#4

1.kdk 2. fh

Unit 3 – HW#4

1.kdk 2. fh

Unit 3 – HW#4

1.kdk 2. fh

Unit 4 – HW#1 answers updated 2018

Unit 4 HW#1 – Part I

1. 65.38 amu 2. 84.93 amu 3. 14.01 amu 4. 17.04 amu

5. 74.44 amu 6. 20.18 amu 7. 58.45 amu 8. 238.03 amu

9. 63.02 amu 10. 107.87 amu

Unit 4: HW#1 – Part II

1. 98.08 g/mol 2. 101.11 g/mol 3. 133.33 g/mol 4. 95.61 g/mol

5. 416.59 g/mol 6. 44.01 g/mol 7. 153.81 g/mol 8. 31.07 g/mol

9. 78.12 g/mol 10. 44.11 g/mol

Unit 4: HW#1 – Part III

1. 6.94 g 2. 26.98 g 3. 60.12 g 4. 12.01 g

5. 215.74 g 6. 601.77 g 7. 270.09 g 8. 250.33 g

9. 1.65 E 10 g 10. 318.29 g

Unit 4: HW#1 – part IV

1. 1 mol Ne 2. 37.38 mol Zn 3. 8.31 x 10-23 mol Ba(OH)2

4. 0.98 mol Ni 5. 0.66 mol Ca 6. 0.5 mol Na 7. 0.047 mol Fe

8. 2.3 x 10-4 mol O2 9. 1347 mol Pb(OH)2 10. 1.41 x 10-13 mol O2

Unit 4: HW#1 – Part V

1. 19.0 g F2 2. 6.06 g Mg 3. 2.6 x 10-10 g Cl 4. 3.7 x 10-3 g Ba(NO3)2

5. 9.6 x 10-21 g WO3

Unit 4: - Part VI

1. 100 g NaOH 2. 94.53 g HNO3 3. 176.2 g SO2 4. 4144 g Pb

5. 9.70 g B 6. 4.3 g S 7. 58.7 g K 8. 4.97 g Pt

9. 1.97 x 10-8 g Au 10. 5000 g (NH4)3PO4

Unit 4: - Part VII

1. 3.27 x 1023 atoms 2. 85.2 g NH3 3. 2 mol H 4. 294.24 g

5. 0.23 mol Zn 6. 66.64 g Hg 7. 1.36E25 atoms 8. 10.0 mol H2O

9. 0.263 mol S 10. 0.173 mol Al 11. 9.09 g H 12. 2000 g H2O

13. 222 mol H 14. 1.34 x 1026 atoms

Unit 3 – HW#4

1.kdk 2. fh

1. 120 miles 2. 1.15 hrs 3a. 200 mph 3b. 400 miles

4a. 8.0 m 4b. 36 m 5a. 200 miles 5b. 0.063 hrs (3.8 min)

6a. 300 s; 5.0 min 6b. 7 cm 7a. 16.7 m/min 7b. 0.28 m/s

7c. 1.00 km/hr 7d. 0.62 mi/hr 8a. 120 km/hr 8b. 33 m/s

8c. 75 mi/hr 9a. 216,000 m 9b. 216 km 10a. 20 s

10b. 0.33 min 10c. 0.0056 hrs

Unit 3 – HW#3

1.