AP CHEMISTRY
SUMMER PACKET
2023
MR. PRIDGEN

Name:

____

AP Chemistry Name: ____________________________

Summer Packet Bishop England High School Summer 2023

Dear Brave Soul,

I pray that you are having an enjoyable and restful summer vacation thus far. Thank you so much for accepting the challenge of AP Chemistry this fall. It will be a fun and exciting year and this summer assignment will help you prepare. We will cover the first three review chapters of the AP Chemistry course work in these packets so that we will have more in class time to prepare for the exam in May. In addition to these first three chapters, you will have to relearn (or learn) your element names and symbols, polyatomic ions, and nomenclature. We will have a quiz on the second full day of class, 8.16.2023, covering nomenclature. You will also have a more in-depth test at the beginning of the second week of school, 8.21.2023, covering the rest of the information from this packet (sig figs, SI units, chemical math, redox, etc.).

Included on the following page is a table of contents that divides the packet into sections. Most of the sections include tutorials, followed by practice questions. If you feel comfortable with the material, do not feel that you need to read the tutorial part of the section. You are required to do every problem and complete all of the material within this packet, which includes they Key Questions and the Exercises for each section. Section X is extra practice and will not be graded but you will be expected to know all of this information. Section XIII is required but will only be graded for completion. This packet should take you between 4 and 7 hours depending on your comfort with chemistry and your retention from your previous chemistry course. It will be due on the orientation day (8.11.2023) and will be graded for accuracy (I will have binder clips for each of you to turn in the packet). **Please keep track of how long each section took you. At the end of the entire packet is a survey that I would like for you to fill out when you finish.**

If you have any difficulties, visit http://bit.ly/PridgenAPChemVideos (you must type it in exactly the same) for some videos that will help you. If you are still struggling, please contact me. I will be checking my email (kpridgen@behs.com) once a week to answer any of your questions. Please resist the temptation to start these at the last minute. These assignments do not require a text, but you should use your sophomore chemistry book if you need extra guidance while I am not around.

Thanks again and have a wonderful summer!

--

In Christ,

Kyle Pridgen

Kyle Pridgen

Science Department

Bishop England High School

Table of Contents

I. Important Mono and Polyatomic Ions (these are the most common, look at page 43 for ions to memorize)

Name	Symbol	Name	Symbol
Hydrogen ion	H^+	Fluoride ion	F^-
Lithium ion	Li^+	Chloride ion	Cl^-
Sodium ion	Na^+	Bromide ion	Br^-
Potassium ion	K^+	Iodide ion	I^-
Copper (I) ion	Cu^+	Hydroxide ion	OH^-
Silver ion	Ag^+	Bicarbonate ion	HCO_3^-
Hydronium ion	H_3O^+	Hypochlorite ion	OCl^-
Ammonium ion	NH_4^+	Nitrate ion	NO_3^-
Magnesium ion	Mg^{2+}	Nitrite ion	NO_2^-
Calcium ion	Ca^{2+}	Thiocyanate	SCN^-
Barium ion	Ba^{2+}	Oxide ion	O^{2-}
Zinc ion	Zn^{2+}	Sulfide ion	S^{2-}
Cadmium ion	Cd^{2+}	Carbonate ion	CO_3^{2-}
Mercury (II) ion	Hg^{2+}	Sulfate ion	SO_4^{2-}
Copper (II) ion	Cu^{2+}	Sulfite ion	SO_3^{2-}
Iron (II) ion	Fe^{2+}	Phosphide ion	P^{3-}
Lead (II) ion	Pb^{2+}	Phosphate ion	PO_4^{3-}
Iron (III) ion	Fe^{3+}		
Aluminum ion	Al^{3+}		

II. Diatomic Molecules (rarely exist alone in nature)

Hydrogen	H_2
Nitrogen	N_2
Oxygen	O_2
Fluorine	F_2
Chlorine	Cl_2
Bromine	Br_2
Iodine	I_2

If they are alone, they will have the word singlet in front of them

e.g. singlet oxygen = O

III. Significant Figures – Start Time: __________ End Time: __________

<u>Exact vs Inexact</u>

What are significant figures, what do they indicate and how are they used in addition, subtraction, multiplication and division?

There are two kinds of numbers in the world:

Exact Numbers

- There are exactly 12 eggs in a dozen.
- Most people have exactly 10 toes and 10 fingers.
- 1 meter = 100 centimeters
- 1 yard = 36 inches
- 1 dollar = 100 cents
- 1 kilometer = 1000 meters

Inexact Numbers

- Any measured value
- Use of a 10 mL graduate cylinder to measure the volume of a solution might give a volume of 8.81 mL (3 sig figs) or a less precise volume of 8.7 mL with a 100 mL graduated cylinder
- An analytical balance might find the mass of a pencil to be 12.1403 g (6 sig figs), while a centigram balance might find it to weigh 12.13 g (4 sig figs)

The number of digits, i.e. significant figures, reported for a ***measured value*** conveys the quality of the measurement and hence, the quality of the measuring device. It is important to use significant figures correctly when reporting a measurement so that it does not appear to be more (or less) precise than the equipment used to make the measurement allows. We can achieve this by controlling the number of digits, or **significant figures**, used to report the measurement.

In this course and in others, *<u>you **must** use correct significant figures in reporting your results</u>*. Laboratory measuring instruments have their limits, just as your senses have their limits. One of your tasks, in addition to learning how to use various measuring instruments properly, will be to correctly determine the precision of the measuring devices that you use and to *<u>report all measured and calculated values to the correct number of significant figures</u>*.

<u>Significant Figure Rules</u>

There are three rules on determining how many significant figures are in a number:

1. **Non-zero digits are always significant.** (see page 4 for details)
2. **Any zeros between two significant digits are significant.** (see page 5 for details)
3. **A final zero or trailing zeros in the decimal portion <u>ONLY</u> are significant.** (see page 5)

Focus on these rules and learn them well. They will be used extensively throughout every Chemistry and Physics course you take in high school and college. You would be well advised to do as many problems as needed to nail the concept of significant figures down tight and then do some more, just to be sure.

<u>Please remember</u> that, in science, with the exception of a few numbers that are defined and hence exact, all numbers are based upon measurements. Since all measurements are uncertain, we must only use those numbers that are meaningful. A common ruler cannot measure something to be 22.4072643 cm long. Not all of the digits have meaning (significance) and, therefore, should not be written down. In science, only the numbers that have significance (derived from measurement) are written.

Rule 1: Non-zero digits are always significant.

Hopefully, this rule seems rather obvious. If you measure something and the device you use (ruler, thermometer, triple-beam balance, etc.) returns a number to you, then you have made a measurement decision and that ACT of measuring gives significance to that particular numeral (or digit) in the overall value you obtain.

Hence a number like 26.38 would have four significant figures and 7.94 would have three. The problem comes with numbers like 0.00980 or 28.09.

Rule 2: Any zeros between two significant digits are significant. (i.e. "sandwiched" zeroes are significant)

Suppose you have a measured value like 406. By the first rule, above, the 4 and the 6 are significant. However, to make a measurement decision on the 4 (in the hundred's place) and the 6 (in the one's place), you HAD to have made a decision on the ten's place. The measurement scale for this number would have calibration marks for the hundreds and tens places with an estimation made in the "ones" place—hence, ***significant figures indicate the number of digits known with certainty*** (e.g. the 1st two digits in 406) ***and one that is an estimate*** (e.g. the 6 in 406). Such a measuring measurement scale would look like this:

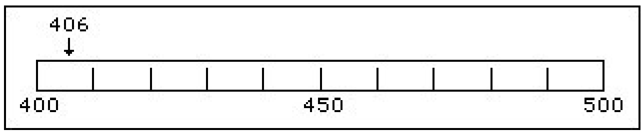


Figure 1. A measuring scale that allows one to use three significant figures

Rule 3: A final zero or trailing zeros in the decimal portion ONLY *are* significant.

This rule causes the most difficulty with students. Here are two examples of this rule with the zeros this rule affects in bold font:

0.005**00**

0.0304**0**

Here are two more examples where the significant zeros are in bold font:

2.3**0** x 10^{-5}

4.5**00** x 10^{12}

Zeros Not Discussed Above

Zero Type #1. Space holding zeros on numbers less than one.

Here are the first two numbers used under rule 3, above. The digits that are NOT significant are underlined:

$0.\underline{00}500 = 5.00 \times 10^{-3}$

$0.\underline{0}3040 = 3.040 \times 10^{-2}$

The underlined zeroes serve only as space holders—their function is to locate the decimal point. They DO NOT involve measurement decisions. The non-significant zeros disappear upon writing the numbers in scientific notation.

Zero Type #2. The zero to the left of the decimal point on numbers less than one.

When a number like 0.00500 is written, the very first zero (to the left of the decimal point) is put there by convention. Its sole function is to communicate unambiguously that the decimal point is a decimal point. If the number were written like this, .00500, there is a possibility that the decimal point might be mistaken for a period. Many students omit that zero. They should not.

Zero Type #3. Trailing zeros in a whole number *without* a decimal point are *not* significant.

200 has only *one* significant figure 25,000 has *two* sig figs

This is based on the way each number is written. When whole number are written as above, the zeros, BY DEFINITION, did not require a measurement decision, thus they are not significant. However, it is entirely possible that 200 really does have two or three or more significant figures. If it does, it will be written differently. Typically, scientific notation, underlining or the use of a decimal point is used for this purpose.

2 significant figures: 2<u>0</u>0 or 2.0×10^2

3 significant figures: 200. or 2.00×10^2

4 significant figures: 200.0 or 2.000×10^2

How will you know how many significant figures are in a number like 200? In a problem without a scientific context, you should be told. If you were doing an experiment, the context of the experiment and its measuring device would tell you how many significant figures to report to people who read the report of your work.

Exact Numbers

Exact numbers, such as the number of people in a room, have an infinite number of significant figures. Exact numbers are counting up how many of something are present, they are not measurements made with instruments. Another example of this are defined numbers, such as 1 foot = 12 inches. There are exactly 12 inches in one foot. Therefore, if a number is exact, it DOES NOT affect the precision of a calculation. <u>Some more examples</u>:

There are 100 years in a century.

2 molecules of hydrogen react with 1 molecule of oxygen to form 2 molecules of water.

1 kilogram = 1000 grams

Each molecule of methane gas, CH_4, contains exactly 1 carbon atom and 4 hydrogen atoms.

Interestingly, the speed of light is now a defined quantity. By definition, the value is 299,792,458 meters per second.

A Brief Aside

There might come an occasion in chemistry when you are not exactly sure how many significant figures are called for. Suppose the textbook mentions 100 mL. You look at this and see only one significant figure. However, an experienced chemist would know that 100 mL can be easily measured to 3 or 4 significant figures. Why then doesn't the textbook (or the professor) write 100.0 (for 4 sig figs) or 1.00×10^2 (for 3 sig figs)?

So, a <u>*brief word of advice*</u>: If you haven't a clue as to how many significant figures to use, try using three or four. These are reasonable numbers of significant figures for most chemical activities.

<u>**Key Questions**</u>

1. What kind of numbers are exact numbers? Give at least one *original* example.

2. What kind of numbers are inexact numbers? Why? Give at least one *original* example.

3. What is your understanding of significant figures: What are significant figures, when should they be used and what function do they serve?

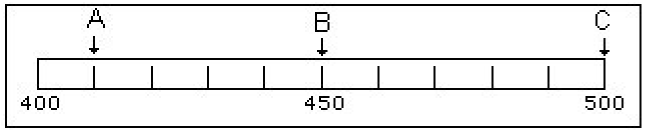


Figure 2. A hypothetical measuring scale

4. What values would you record for measurements A, B and C if each measurement fell on the line each arrow points to in figure 2, above? How many sig figs should each measurement have?

A = ________________ B =________________ C = ________________

5. Later in the quarter you will be asked to measure out accurately about 3 grams of an unknown salt with a milligram electronic balance (a balance that measures out to the nearest milligram, 0.001 g). What mass of salt should you measure out? How many significant figures should you record?

6. Suppose you are asked to measure out about 25 mL of deionized water *as accurately as you can*.

a.) What measuring device would you use? ____________________________

b.) How much water should you measure out? ___________________________

c.) How many significant figures would you report? _____________________

Exercises

7. How many significant figures are there in each of the following numbers? Record your responses in the spaces provided and circle the digits that are significant.

a.)	3.0800	____________	f.)	3.200×10^9	____________
b.)	0.00418	____________	g.)	250	____________
c.)	7.09×10^5	____________	h.)	780,000,000	____________
d.)	91,600	____________	i.)	0.0101	____________
e.)	0.003005	____________	j.)	0.00800	____________

Rounding Numbers

In numerical problems, it is often necessary to round numbers to the appropriate number of significant figures. Consider the following examples in which each number is rounded so that each of them contains 4 significant figures. Study each example and make sure you understand why they were rounded as they were:

Original number	→	Number rounded to 4 sig figs
41,008	→	41,010
1.25624	→	1.256
0.017837	→	0.01784
120	→	120.0
127.450	→	127.4
127.4501	→	127.5
127.550	→	127.6
127.25000	→	127.2
127.25001	→	127.3
127.35	→	127.4

Key Questions

8. Summarize the rounding rule(s) used in the ***first three examples***, above.

9. Summarize the rounding rule(s) used in the last six examples (i.e. those using 127), above. This rule is often referred to as the "odd – even" rule. (Hint: look at the value of the tenths place)

Exercises

10. Round the following numbers to four significant figures.
   a.) 2.16347×10^5 =
   d.) 7.2518 =
   b.) 4.000574×10^6 =
   e.) 375.6523 =
   c.) 3.6825 =
   f.) 21.865001 =

11. Round off each number to the indicated number of significant figures (sf).
   a.) 231.554 (to 2 sf) =
   e.) 249,441 (to 3 sf) =
   b.) 0.00845 (to 2 sf) =
   f.) 0.00250122 (to 3 sf) =

c.) 150,000 (to 1 sf) = g.) 12,049,002 (to 4 sf) =

d.) 0.0023 (to 3 sf) = h.) 0.00200210 (to 3 sf) =

<u>Using Significant Figures in Addition and Subtraction</u>

Did you know that 30,000 plus 1 does not always equal 30,001? In fact, 30,000 + 1 is sometimes equal to 30,000! You may find this hard to believe, but let's examine this.

Recall that zeros in a number are not always *significant*. Knowing this makes a big difference in how we add and subtract. For example, consider a swimming pool that can hold 30,000 gallons of water. If I fill the pool to the maximum fill line and then go and fill an empty one-gallon milk jug with water and add it to the pool, do I then have exactly 30,001 gallons of water in the pool? Of course not. I had approximately 30,000 gallons before and after I added the additional gallon because "30,000 gallons" is not a very precise measurement. So, we see that sometimes 30,000 + 1 = 30,000!

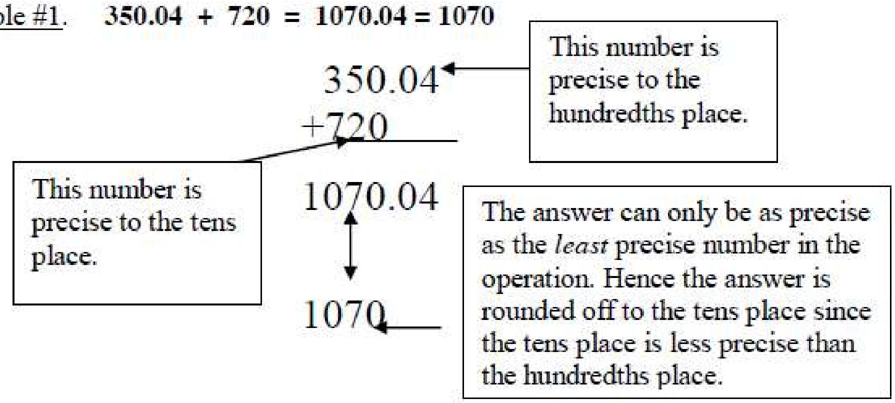
In mathematical operations involving significant figures, the answer is reported in a way that reflects the reliability of the *<u>least precise</u>* number. ***An answer is no more precise that the least precise number used to get the answer***. Imagine a team race where you and your teammates must finish together at the same time. Who dictates the speed of the team? Of course, the slowest member of the team. Your answer cannot be MORE precise than the least precise measurement.

Use the "decimal rule" when adding and subtracting numbers:

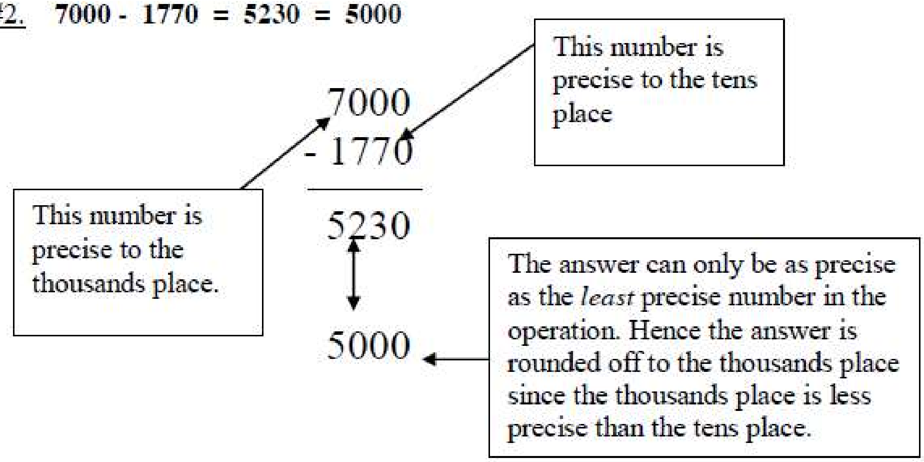
For addition or subtraction, the answer must be rounded off to contain only as many decimal places as are in the value with the *<u>least</u>* number of decimal places.

WARNING!! The rules for addition/subtraction are different from those of multiplication/division. A very common student error is to swap the two sets of rules. Another common error is to use just one rule for both types of operations.

<u>Example #1.</u> **350.04 + 720 = 1070.04 = 1070**



<u>Example #2.</u> **7000 - 1770 = 5230 = 5000**



Key Questions

12. Consider example #1 from above. Indicate in the spaces below the number of significant figures (sf) for each number in the problem.

350.04	+	720	=	1070.04	=	1070
____sf		**____sf**		**____sf**		**____sf**

Should the number of significant figures be considered when adding or subtracting measured numbers? *Explain*

13. When you add and subtract numbers, how do you identify the first uncertain number in the result?

Exercises

Record the answer before and after rounding off for each problem below.

14. 3.461728 + 14.91 + 0.980001 + 5.2631 = ________________ = __________________
15. 23.1 + 4.77 + 125.39 + 3.581 = __________________ = _________________________
16. 22.101 - 0.9307 = ___________________ = ________________________
17. 0.04216 - 0.0004134 = ___________________________ = _________________________
18. 564,321 - 264,321 =__________ = __________________

Using Significant Figures in Multiplication and Division

A chain is no stronger than its weakest link—that is, an answer is no more precise that the least precise number used to get the answer.

Use the "Chain Rule" when multiplying and dividing measured numbers:

When measurements are multiplied or divided, the answer can contain no more significant figures than the number with the fewest number of significant figures. This means you MUST know how to recognize significant figures in order to use this rule.

To round correctly, follow these simple steps:

1) Count the number of significant figures in each number.
2) Round your answer to the least number of significant figures.

Example #1.

$$\frac{4560}{14} = 325.714285714 = 330$$

3 sig figs

2 sig figs

The rounded answer has only 2 significant figures since 2 is the least number of significant figures in this problem.

Example #2.

$$13.1 \times 1.2039 = 15.77109 = 15.8$$

3 sig figs

5 sig figs

The rounded answer has 3 significant figures since 3 is the least number of significant figures in this problem.

Multi-Step Calculations: Keep at Least One Extra Significant Figure in Intermediate Answers

When doing multi-step calculations, *keep at least one more significant figure in intermediate results* than needed in your final answer. For example, if a final answer requires two significant figures, then carry *at least* three significant figures in all calculations. If you round-off all your intermediate answers to only two digits, you are discarding the information contained in the third digit, and as a result the *second* digit in your final answer might be incorrect. This phenomenon is known as "roundoff error." <u>*Avoid rounding errors by carrying at least on extra sig fig throughout a multi-step calculation and then round off to the correct number of sig figs at the very end.*</u>

<u>Key Questions</u>

19. When you multiply and divide numbers, what is the relationship between the number of significant figures in the result and the number of significant figures in the numbers you are multiplying or dividing?

Exercises

Record the answer <u>*before*</u> and <u>*after*</u> rounding off for each problem below.

20. $(3.4617 \times 10^{7}) \div (5.61 \times 10^{4})$ = ________________________ = __________________________

21. $[(9.714 \times 10^{5})(2.1482 \times 10^{9})] \div [(4.1212)(3.7792 \times 10^{5})]$ = __________________________

(Watch your order of operations on this problem!) = _________________________

22. $(4.7620 \times 10^{15}) \div [(3.8529 \times 10^{12})(2.813 \times 10^{7})(9.50)]$ = __________________________

= __________________________

23. $[(561.0)(34{,}908)(23.0)] \div [(21.888)(75.2)(120.00)]$ = __________________________

= __________________________

24. Carry out each of the following calculations. *Check that each answer has the correct number of significant figures and the correct units of measure*.

a.) $\dfrac{2.420\ g\ +\ 15.6\ g}{4.8\ g} =$

b.) $\dfrac{7.87\ g}{16.1\ mL\ -\ 8.44\ mL} =$

c.) $V = \pi r^2 h \quad where\ r = 6.23\ cm$ and $h = 4.630\ cm$

V =

d.) $\dfrac{8.32\ x\ 10^7 g}{\frac{4}{3}(3.1416)(1.95\ x\ 10^2 cm)^3} =$

Note: 4/3 is an exact number!

e.) $E_k = \frac{1}{2}mv^2 = \dfrac{(1.84\ x\ 10^2 g)(44.7\ {}^{m}/_{s})^2}{2} =$

f.) $\dfrac{(1.07\ x\ 10^{-4}\frac{mol}{L})^2\ (2.6\ x\ 10^{-3}\frac{mol}{L})}{(8.35\ x\ 10^{-5}\frac{mol}{L})(1.48\ x\ 10^{-2}\frac{mol}{L})^3} =$

IV. *Système international (d'unités)* – Start Time: __________ End Time: __________

A series of international conferences on weights and measures has been held periodically since 1875 to refine the metric system. At the 11th conference held in France in 1960, a new system of units known as the ***International System of Units*** (abbreviated SI in all languages) was proposed as a replacement to the metric system. The five of the seven base units for the SI system are given in **Table 1—**missing from the table are electric current (ampere) and luminous intensity (candela), units that aren't of interest to us right now. You will need *to memorize the base unit and its symbol for each of Tables 1 & 2* as they will be used extensively this year in AP Chem.

Table 1. SI base units and their symbols

Physical Quantity	Name of base unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	Kelvin	K
Amount of substance	mole	mol

Derived SI Units

The units of every measurement in the SI system, no matter how simple or complex should be derived from one or more of the seven base units. For example, the preferred unit for volume is the cubic meter, m^3, because volume has units of length cubed and the SI unit for length is the meter. Strict adherence to SI units would require changing directions such as "add 250 mL of water to a 1-Liter beaker" to add 0.00025 cubic meters of water to an 0.001 m^3 container. Because of this, a number of units that are not strictly acceptable under the SI convention are still in use. Some of the common non-SI units in common in use are in Table 2.

Table 2. Non-SI units in common use.

Physical Quantity	Name of base unit	Symbol
Volume	liter	L
Temperature	degree Celsius, Kelvin	^{o}C, K
Concentration	molarity	*M*
Pressure	Atmosphere, torr = mmHg	atm, torr, mmHg

Table 3. Common decimal prefixes used with SI units of measure that *should be memorized*

Prefix	Prefix Symbol	Meaning	Exponential Notation
Mega	M	million	1×10^{6}
Kilo	k	thousand	1×10^{3}
Centi	c	hundredth	1×10^{-2}
Milli	m	thousandth	1×10^{-3}
Micro	µ	millionth	1×10^{-6}
Nano	n	billionth	1×10^{-9}

<u>Note the following examples</u>:

- "milli" means one-thousandth; so a milliliter (symbol: mL) is one thousandth of a Liter and it takes 1000 mL to make 1 L; therefore 1 L = 1000 mL
- "kilo" means one thousand; so "kilogram" (kg) means one thousand grams: 1 kg = 1000g. One millionth of a gram would be represented by one microgram (µg): 1 Pg = 10^{-6} g
  It takes one million micrograms to equal one gram: 10^{6} µg = 1 g
- one centimeter (1 cm) is equal to 0.01 m because one cm is "one hundredth of a meter":
  1 cm = 0.01 m = 10^{-2} m

<u>Key Questions</u>

1. How many milligrams are there in one kilogram?

2. How many meters are in 21.5 km?

3. Is it possible to answer this question: How many mg are in one km? *<u>Explain</u>*.

4. a.) What is the difference between a Mm and a mm?

   b.) Which is larger 1.0 Mm or 1.0 mm? *<u>Explain.</u>*

5. Complete the table below by indicating what physical quantity each measurement is measuring.

Measurement	**Physical Quantity Measured**
1.33 m^3	
298.6 K	
3.47 L	
44 µg	
8.75 mm	
760 torr	

Measurement	**Physical Quantity Measured**
22.8 °C	
2.0 *M*	
3.8 atm	
78.4 mL	
18.3 cm	
44.8 ng	

The Factor-Label Method

In scientific measurements, units usually follow the numerical value. When mathematically manipulating scientific measurements, you will often need to convert from one unit to another. When this is done, you must multiply what is given by one. In the ***Factor-Label Method*** (also called "*dimensional analysis*"), we multiply what is given by one or more conversion factors. In a conversion factor, what is in the numerator must be equivalent to what is in the denominator (otherwise, you're not multiplying by one!). You must apply conversion factors such that the units cancel appropriately.

Therefore, you will constantly be asking the questions: "What units do I have? Is it in the numerator or in the denominator? What units can I convert to from what I have?"

Units you have but no longer want will be written in the conversion factor on the opposite side of the division bar.

For example, if the units you have but want to convert from are in the numerator, those units will be put in the denominator of the conversion factor.

Example 1. Using more than one conversion factor

By definition, 1 inch is equivalent to 2.54 centimeters, 1 foot is equivalent to 12 inches, and 1 mile is equivalent to 5280 feet. How many kilometers are in a mile?

$$1\text{ mi} \times \frac{5280\text{ ft}}{1\text{ mi}} \times \frac{12\text{ in}}{1\text{ ft}} \times \underbrace{\frac{2.54\text{ cm}}{1\text{ in}} \times \frac{1\text{ m}}{100\text{ cm}} \times \frac{1\text{ km}}{1000\text{ m}}}_{etc.} = 1.609344\text{ km}$$

What we've started with: the distance is 1 mile.

What can we convert from miles to? In our list of givens, we are told that "1 mile is equivalent to 5280 feet." Since we have "mi" in the numerator, "mi" must be placed in the denominator in the conversion factor. Consequently, "ft" goes in the numerator. Since "1 mile is equivalent to 5280 feet," the number "5280" accompanies "ft" and "1" goes with "mi".

What can we convert from feet to? In our list of givens, we are told that "1 foot is equivalent to 12 inches." Since we have "ft" in the numerator, "ft" must be placed in the denominator in the conversion factor. Consequently, "in" goes in the numerator. Since "1 foot is equivalent to 12 inches," the number "12" accompanies "in" and "1" goes with "ft".

etc.

Example 2. Conversion factor raised to a power

The density of isopropyl alcohol is 6.56 pounds per gallon. What is the density of isopropyl alcohol in grams per cubic centimeter? (1 kg = 2.205 lb, 1 gal = 3.785 L, 1 L = 1 dm^3)

$$\frac{6.56\text{ lb}}{1\text{ gal}} \times \frac{1\text{ kg}}{2.205\text{ lb}} \times \underbrace{\frac{1000\text{ g}}{1\text{ kg}} \times \frac{1\text{ gal}}{3.785\text{ L}} \times \frac{1\text{ L}}{1\text{ dm}^3}}_{etc.} \times \left(\frac{1\text{ dm}}{10\text{ cm}}\right)^3 = 0.786\text{ g/cm}^3$$

What we've started with: the density is 6.56 pounds per gallon.

What can we convert from pounds to? In our list of givens, we are told that "1 kg = 2.205 lb." Since we have "lb" in the numerator, "lb" must be placed in the denominator in the conversion factor. Consequently, "kg" goes in the numerator. Since "1 kg = 2.205 lb," the number "1" accompanies "kg" and "2.205" goes with "kg".

etc.

Before this conversion factor, the units are g/dm^3. So far "dm" is in the denominator. Therefore, in this conversion factor, "dm" must be in the numerator. Consequently, "cm" goes in the denominator. We know from the prefixes "deci" and "centi" that there are 10 centimeters in 1 decimeter, so the numbers of the conversion factor are "1" in the numerator and "10" in the denominator. But since the units of dm are cubed, we must cube dm in the conversion factor. To maintain that what is in the numerator is equal to what is in the denominator, the entire conversion factor is cubed.

<u>Key Questions</u>

6. What in the numerator of a conversion factor must be equivalent to what is in the denominator of the conversion factor. Explain why.

7. Consider the exercise: "How many seconds are there in 50 minutes?" Suppose it was solved in the following way:

$$50.\text{ min} \times \frac{1\ min}{60\ s} = 0.83\ s$$

a.) Is the answer reasonable? ***<u>Explain.</u>***

b.) Comment on how the units were handled. What "rule" of the model (the factor-label method) was broken in the above solution?

Exercises

Instructions*:* Show your work using the factor-label method. *Circle your answers* and use correct units and significant figures for all questions—no work, no credit. Use the table of SI – English conversion factors on the inside back cover of your text as needed.

8. Perform the following unit conversions Show your work in each case using the factor-label method. *Circle your answers and use correct sig figs.*

   a.) 15.60 cm = ?? m

   b.) 41.0 kg = ?? µg

   c.) 9.2 mL = ?? µL

   d.) 9.16×10^{-5} m = ?? nm

9. The distance between two adjacent peaks on a wave is called the wavelength. The wavelength of visible light determines its color.

   a.) The wavelength of a beam of green light is 545 nm. What is its wavelength in meters?

   b.) The wavelength of a beam of red light is 683nm. What is its wavelength in inches?

10. A home fashions designer from Europe comes to the United States and decides to purchase some gorgeous fabric that she knows she cannot find back in her home town. She asks the salesperson at the fabric store to cut 36 m^2 of the material. The salesperson unfortunately does not have a measuring tape that gives meters, but she has one that measures yards (1 yd = 3 ft). How many square yards does the salesperson need to cut?

11. One manufacturer boasts that their car offers a gas mileage of 32 mi/gal. A European manufacturer advertises that their car has a gas mileage of 15 km/L. Which car would be more economical to operate on the basis of gas mileage? *Justify your answer with the appropriate calculation(s).* (1 gal = 3.785 L)

12. The volume of a certain bacterial cell is 1.72 μm^3.

a.) What is the cell's volume in cubic millimeters, mm^3?

b.) What is the volume in liters of 1.0×10^5 of these bacterial cells?

13. An Olympic-size pool is 50.0 m long and 25.0 m wide.

a.) How many gallons of water (assume d_{water} = 1.00 g/mL) are needed to fill the pool to an average depth of 4.8 ft?

b.) What is the mass in kg of water in the pool?

14. Nutritional tables give the potassium content of a standard apple (about 3 apples per pound, lb.) as 159 mg. How many grams of potassium are in 4.25 kg of apples? (1 lb = 453.592 g)

V. Ionic Compounds and Acids – Start Time: __________ End Time: __________

Compounds Containing Polyatomic Ions

In addition to monatomic ions, there are polyatomic ions. Within a **polyatomic ion**, each atom is connected to one or more atoms of the same polyatomic ion. Such connected atoms are **covalently bound** to each other. Atoms connected by a covalent bond share electrons with each other (as opposed to neighboring cations and anions, in which electrons had been transferred from the former to the latter species).

Table 1. Common polyatomic ions. See page 43 for a table of the ions that you need to know in this class.

Cations:	NH_4^+	Ammonium	Hg_2^{2+}	Mercury (I)
	H_3O^+	Hydronium		
Anions:	$C_2H_3O^{2-}$	Acetate	OH^-	Hydroxide
	NH_2^-	Amide	ClO^-	Hypochlorite
	N^{3-}	Azide	NO_3^-	Nitrate
	HCO_3^-	Hydrogen carbonate	$C_2O_4^{2-}$	Oxalate
		(*or bicarbonate*)	O_2^{2-}	Peroxide
	BO_3^{3-}	Borate	MnO_4^-	Permanganate
	CO_3^{2-}	Carbonate	ClO_4^-	Perchlorate
	C_2^{2-}	Carbide	PO_4^{3-}	Phosphate
	ClO_3^-	Chlorate	SiO_4^{4-}	Silicate
	CrO_4^{2-}	Chromate	SO_4^{2-}	Sulfate
	OCN^-	Cyanate	O_2^-	Superoxide
	CN^-	Cyanide	$S_2O_3^{2-}$	Thiosulfate
	$Cr_2O_7^{2-}$	Dichromate	I_3^-	Triiodide

Compounds composed of polyatomic ions form crystals with structures similar to that for compounds of monatomic ions – cations and anions stack on top of each other in a periodic array. Suppose we looked into a crystal of Hg_2I_2 , mercury (I) iodide. We would find planes containing the clusters of ions as shown on the right.

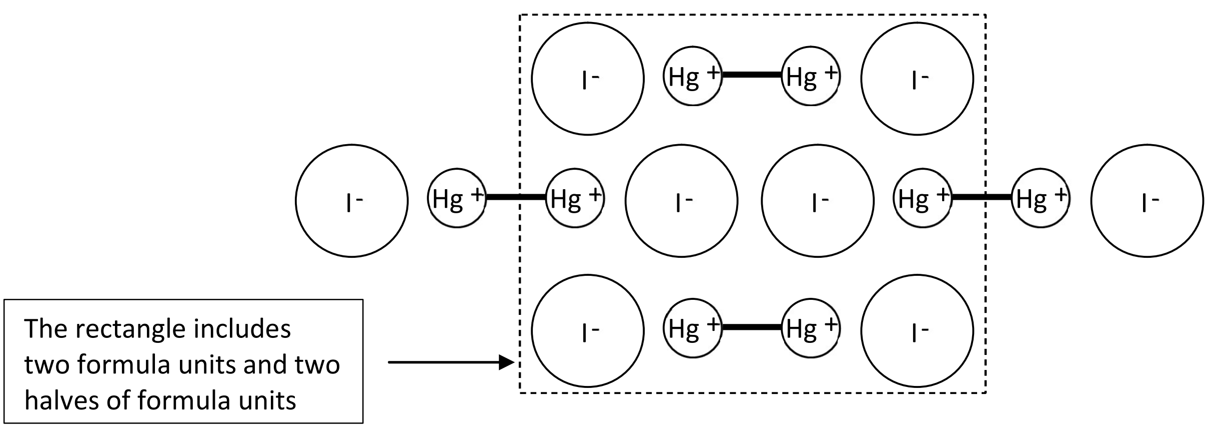


Figure 1. A section of a crystal of mercury (I) iodide, Hg_2I_2

Key Questions

1. What kinds of bonding exist in mercury (I) iodide?

2. Why is the formula for mercury (I) iodide "Hg_2I_2" and not simply "HgI"?

Nomenclature of Ionic Compounds

Formula	Unambiguous Name	Formula	Unambiguous Name
LiF	lithium fluoride	$Mn_3(PO_4)_2$	manganese(II) phosphate
Na_2O	sodium oxide	$MnPO_4$	manganese(III) phosphate
ZnS	zinc sulfide	$Cu(IO_3)_2$	copper(II) iodate
$Al(NO_3)_3$	aluminum nitrate	SnI_2	tin(II) iodide
NH_4NO_2	ammonium nitrite	Hg_3N_2	mercury(II) nitride
$MgSO_3$	magnesium sulfite	Hg_2SO_4	mercury(I) sulfate
K_3AsO_4	potassium arsenate	$PbSeO_4$	lead(II) selenate

Key Questions

3. The name of an ionic compound typically consists of two words separated by a space. Is the first word in the name of an ionic compound that of the cation or that of the anion? Is the second word that of the cation or that of the anion?

4. Look at the examples in the Model, paying attention to the differences between compounds with anions having similar names (*e.g.*, ***nitride*** versus ***nitrate*** versus ***nitrite***; ***sulfide*** versus ***sulfate*** versus ***sulfite***; ***iodide*** versus ***iodate***).

   a. What is the difference between an anion whose name ends in ***-ide*** and an anion whose name ends in ***-ate*** (or ***-ite*** for that matter)? (*i.e.*, What does an ***-ate*** anion have that an ***-ide*** anion does not?)

   b. What is the charge (both magnitude and sign) of the ***nitrate anion***? Based on the formula of ***ammonium nitrite***, what is the charge of the nitrite anion?

c. What is the charge (both magnitude and sign) of the ***sulfate anion***? Based on the formula of ***magnesium sulfite***, what is the charge of the ***sulfite anion***?

d. In general, if you've memorized the formulas of polyatomic anions with names ending in ***-ate***, you need not memorize the formulas of the corresponding polyatomic anions ending in ***-ite***. In going from ***-ate*** to ***-ite*** (*e.g.*, from nitrate to nitrite, or from sulfate to sulfite), what happens to the formula (both in terms of the number of each type of atom in the ion and the overall charge)?

e. There is a polyatomic ion called "phosphite". Given that the formula of the phosphate ion is PO_4^{3-}, what is the formula (including the charge) of the phosphite ion?

5. Phosphorous and arsenic are in the same column on the Periodic Table. Elements in the same "family" or "Group" will very often have similar chemical reactivities. Since there is a polyatomic ion known as "phosphate", it should not be surprising that there is also an "arsenate". What is the formula for the polyatomic anion called "tellurite"? *Hint:* tellurium, Te, is in the same group on the periodic table, group VIA, as sulfur.

6. LiF is not "lithium (I) fluoride" and it would be ambiguous to refer to $Mn_3(PO_4)_2$ as "manganese phosphate".
   a. What does the Roman numeral in parentheses after the name of a metal tell the reader?

   b. When should the Roman numeral in parentheses after the name of a metal be written?

7. a. What identifies the compound $Sb_2(SO_4)_3$ as an ionic compound?

   b. Keeping in mind that antimony cations could have more than one possible charge, what is the unambiguous name of $Sb_2(SO_4)_3$?

Nomenclature of Acids

Formula	Unambiguous Name	Formula	Unambiguous Name
HCl	hydrochloric acid	H_2SO_4	sulfuric acid
H_2S	hydrosulfuric acid	H_2SO_3	sulfurous acid
HCN	hydrocyanic acid	$HClO_4$	perchloric acid
		$HClO_3$	chloric acid
		$HClO_2$	chlorous acid
		HClO	hypochlorous acid

Key Questions

8. What are the names of the anions in *hydro*chlor*ic* acid, *hydro*sulfur*ic* acid, and *hydro*cyan*ic* acid?

   Hydrochloric acid ______________________________

   Hydrosulfuric acid ______________________________

   Hydrocyanic acid ______________________________

   b. What common suffix do these anions share?

9. a. What are the names of the anions in sulfur*ic* acid and chlor*ic* acid?

   Sulfuric acid _____________________ Chloric Acid _____________________

   b. What common suffix do these anions share?

10. a. What are the names of the anions in sulfur*ous* acid and chlor*ous* acid?

    Sulfurous acid _____________________ Chlorous Acid _____________________

    b. What common suffix do these anions share?

11. Refer back to Key Questions 8, 9, and 10. In your own words, summarize the rules for…

    a. when the ***hydro***- prefix and ***-ic*** suffix is used when naming an acid.

b. when *only* the ***-ic*** suffix is used in naming an acid.

c. when *only* the ***-ous*** suffix is used in naming an acid.

12. What is the name of the polyatomic anion in each of the following acids?

    a. *Per*chlor*ic* acid ___________________________

    b. *Hypo*chlor*ous* acid ________________________

Elements in the same "family" (*i.e.*, group or column of the Periodic Table) will very often have similar chemical reactivities. Since there is an acid known as "chloric acid", it should not be surprising that there is also be a "bromic acid". Write the formulas for the following acids.

a. perbromic acid ____________________ c. bromous acid ____________________

b. bromic acid ____________________ d. hypobromous acid ________________

Exercises

Provide the correct, unambiguous name for each of the following. Acids are indicated if followed by (aq), meaning in aqueous solution—i.e. dissolved in water.

13. HgS
14. $Al_2(CO_3)_3$
15. H_3PO_4 (aq)
16. BaO
17. $K_2Cr_2O_7$

18. H_2SO_4 (aq)

19 HCl (g)

20. $HClO$ (aq)

21. $Ca(HCO_3)_2$

22. $Pb(NO_3)_2$

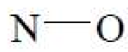
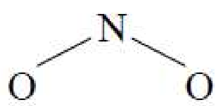
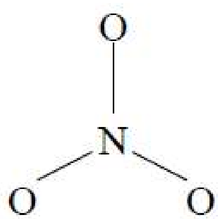
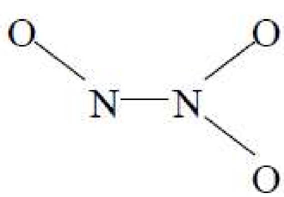
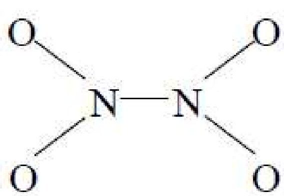
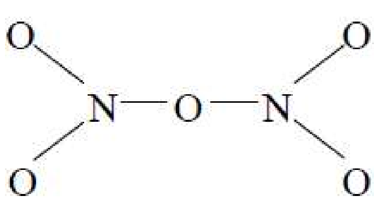
<u>Provide the correct formula for each of the following</u>:

23. Mercury (I) nitrite
24. Iron (III) nitride
25. Strontium sulfite
26. Manganese (II) phosphite
27. Sodium dihydrogen phosphate
28. Phosphorous acid
29. Copper (I) hydrogen phosphate
30. Hydrofluoric acid

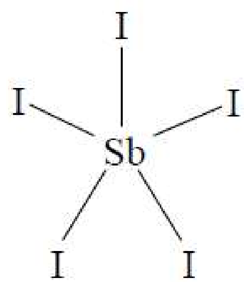
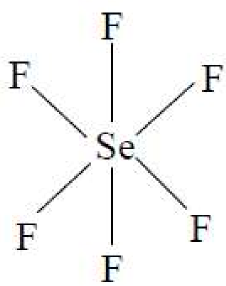
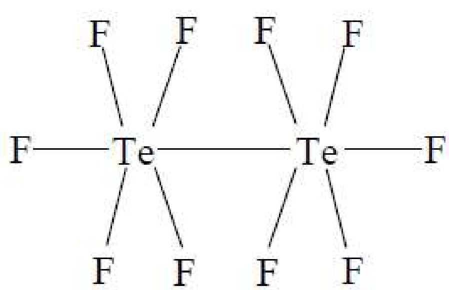
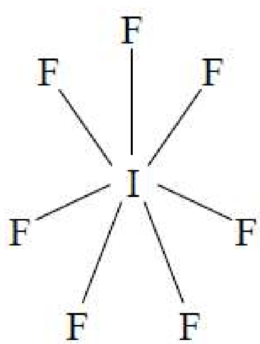
VI. Covalent Compounds – Start Time: __________ End Time: __________

<u>Binary Molecular Nomenclature</u>

There are several compounds that are all *oxides of nitrogen*. They are:

Representation	Formula	Unambiguous Name
	**NO**	**nitrogen monoxide**
	**NO_2**	**nitrogen dioxide**
	**N_2O**	**dinitrogen monoxide**
	**NO_3**	**nitrogen trioxide**
	**N_2O_3**	**dinitrogen trioxide**
	**N_2O_4**	**dinitrogen tetroxide** (or dinitrogen tetraoxide)
	**N_2O_5**	**dinitrogen pentoxide** (or dinitrogen pentaoxide)

The atoms around a central atom repel each other, so there is an upper limit to how many bonds a central atom may have. As the central atom becomes larger, it allows more atoms to surround it without these atoms repelling each other. A number of examples of molecules having several covalent bonds are:

Representation	Formula	Unambiguous Name
	SbI_5	antimony pentaiodide
	SeF_6	selenium hexafluoride
	Te_2F_{10}	ditellurium decafluoride
	IF_7	iodine heptafluoride

Key Questions

1. From the Model, what is meant by "***binary molecular compound***"?

2. Classify the elements in the compounds presented in the Model as metals (***M***), nonmetals (***N***), or semimetals (metalloids) (***S***). The elements below are presented in order of increasing atomic number. *Circle your answers below*.

Nitrogen: ***M*** or ***N*** or ***S*** ?	Selenium: ***M*** or ***N*** or ***S*** ?
Oxygen: ***M*** or ***N*** or ***S*** ?	Antimony: ***M*** or ***N*** or ***S*** ?
Fluorine: ***M*** or ***N*** or ***S*** ?	Tellurium: ***M*** or ***N*** or ***S*** ?
Sulfur: ***M*** or ***N*** or ***S*** ?	Iodine: ***M*** or ***N*** or ***S*** ?

b. When elements belonging to the classifications you listed in (a) combine, what type of compound are they likely to form: molecular or ionic? (You may circle your answer.)

c. How can you tell from the formula of a compound if it is a binary molecular compound? Hint: What kind of elements are in a binary molecular compound.

3. a. How do the prefixes “mono-”, “di-”, “tri-”, *etc*. in the names in the Model help the reader?

   b. Pertaining specifically to the family of nitrogen oxides, explain why it is so important to use the prefixes when naming a compound?

4. From the Model, what is the apparent rule for using the prefix “mono-”? When is it not used?

5. With what suffix does the name of a binary compound always end?

6. For each of the binary compounds presented in the Model, find the relative positions of the two elements on the Periodic Table. Based on the positions of any two nonmetals/metalloids on the Periodic Table, state a general rule that is used to determine which element’s name is written first in the compound’s name. For example, NO is “nitrogen monoxide”. Its formula is not “ON” nor is it called “oxygen mononitride”. Why?

Exercises

Provide the correct unambiguous name for each of the following binary molecular compounds.

7. Br_3O_8 ___________________________________

8. I_4O_9 ___________________________________

9. ICl ___________________________________

10. AsF_5 ___________________________________

11. CO_2 ___________________________________

Write the correct formulas for each of the following binary molecular compounds:

12. dichlorine monoxide ____________________

13. sulfur trioxide ____________________

14. tetraphosphorous heptasulfide ________________

15. disilicon hexaiodide ____________________

16. selenium tetrabromide ____________________

17. bromine pentafluoride ____________________

18. Like the nitrogen oxides, there is a "family" of sulfur fluorides: S_2F_2, SF_4, SF_6, and S_2F_{10}. Attempt to sketch a representation of each one (use those in the Model as a guide) and next to each representation provide the unambiguous name for the four sulfur fluorides.

a. S_2F_2	c. SF_6
b. SF_4	d. S_2F_{10}

VII. Atomic Mass and Moles – Start Time: __________ End Time: __________

Average Atomic Mass

A single atom is extremely small. The typical atom will have a mass of approximately 3 x 10^{-23} g. The smallest mass that the standard analytical balance can weigh reliably is 0.0001 g, which corresponds to roughly 3 quintillion (*i.e.*, 3,000,000,000,000,000,000) atoms. Therefore, we define

$$1 \text{ amu} = 1.6606 \times 10^{-24} \text{ g}$$

to make it convenient to discuss the mass of very small particles in terms of **atomic mass units** (amu) rather than very tiny fractions of what can be weighed out on a balance.

An overwhelming majority of the elements that are encountered in the chemistry lab have two or more naturally-occurring isotopes. If an element has more than one naturally-occurring isotope, then a random sample of the element should be assumed to exist as a mixture of these isotopes that are found in Nature.

Carbon has been found to be 98.89% ^{12}C and 1.11% ^{13}C. Carbon-12 (chosen by the scientific community to define the amu) has an **isotopic mass** of exactly 12 amu (*i.e.*, one ^{12}C atom weighs 12 amu) and that of ^{13}C is 13.0034 amu.

It is assumed that the composition of a sample of an element (in terms of the **percent natural abundances** of each of the element's isotopes) is the same everywhere on Earth. Therefore, in any sample containing carbon (be it a diamond or an organic compound containing carbon in addition to other elements), 98.89% of the carbon atoms in that sample will be C-12 (each weighing 12 amu) and the remaining 1.11% of the carbon atoms will be C-13 (each weighing 13.0034 amu), regardless of where the sample was taken from.

We now determine the average atomic mass of carbon in a manner that is organized and easy to follow (*i.e.*, by tabulating what we know and what we can derive from our given information):

Isotope	Isotopic Mass (amu)	Percent Natural Abundance	Mass Contribution (amu)
^{12}C	12.0000	98.89	(12.0000 x 0.9889) = 11.87
^{13}C	13.0034	1.11	(13.0034 x 0.0111) = 0.144
			Average Mass (amu) (11.87 + 0.144 =) 12.01

Key Questions

1. If you were able to select one carbon atom at random, what is the mass of that atom most likely to be (in amu)? Why?

2. Yes or No (Circle your answer.): Does any carbon atom anywhere in the Universe have a mass equivalent to the average atomic mass of carbon on Earth? *Briefly explain your reasoning*.

3. a. What is the mass in amu of ten sextillion (*i.e.*, 10^{22}) Carbon-12 atoms? *Show your work* and *circle your answer.*

   b. What is the mass in grams of ten sextillion Carbon-12 atoms? *Show your work* and *circle your answer.*

   c. What is the mass of ten sextillion Carbon-13 atoms (in g)? *Show your work* and *circle your answer.*

4. If a diamond consisting of ten sextillion "randomly-selected" carbon atoms (roughly a "one-karat diamond") were placed on an analytical balance, the balance would read (circle your answer and briefly explain your reasoning below):

   a.) slightly less than 0.1993 g
   b.) 0.1993 g
   c.) slightly more than 0.1993 g
   d.) slightly less than 0.2159 g
   e.) 0.2159 g
   f.) slightly more than 0.2159 g

   Explanation:

Exercises

5. Europium has two naturally-occurring isotopes. Use the data in the table below to determine the average atomic mass of europium to the nearest 0.1 amu. *Hint*: Complete the table similarly to what is shown for carbon in the Model.

Isotope	Isotopic Mass (amu)	Percent Natural Abundance	
151Eu	150.92	47.8	
153Eu	152.92	52.2	
		Average atomic mass (amu)	

6. 20.5% of germanium is Ge-70, which has an isotopic mass of 69.924 amu. 27.4% of germanium is Ge-72, which has an isotopic mass of 71.922 amu. 7.8% of germanium is Ge-73, which has an isotopic mass of 72.923 amu. 36.5% of germanium is Ge-74, which has an isotopic mass of 73.921 amu. The rest of naturally-occurring germanium is Ge-76, which has an isotopic mass of 75.921 amu. Use this data to calculate the average atomic mass of germanium to the nearest 0.01 amu. *<u>Hint</u>*: Complete the table below to help you organize the information given.

7. Copper has two naturally occurring isotopes, ^{63}Cu (isotopic mass 62.9396 amu) and ^{65}Cu (isotopic mass 64.9278 amu). If copper has an atomic mass of 63.546 amu, what is the percent abundance of each element? *<u>Show your work</u>* and *<u>circle your answer</u>*.

Key Questions

8. a. Look at the Periodic Table. How do the numbers under the symbols for carbon, europium, and germanium compare to 12.01 amu and the values you determined in Exercises 5 and 6 respectively?

   b. What is the number under the symbol of an element on the Periodic Table?

1 dozen items = 12 items	1 gross of items = 144 items
1 score of items = 20 items	1 **mole** of items = 6.022 u 10^{23} items

Exercises

9. a. One often buys donuts by the dozen. How many donuts are there in a dozen donuts?

   b. Abraham Lincoln started his Gettysburg Address with the words, "Four score and seven years ago…" How many years are there in a score of years?

   c. Cheap plastic jewelry is very often sold "by the gross". How many plastic necklaces are there in a gross of necklaces?

   d. A mole of liquid water occupies a volume of approximately 18 mL. How many water molecules are in a mole of water molecules?

10. a. 1 mol of items is 6.022×10^{23} of those items. 1 carbon atom weighs on average 12.01 amu. 1 amu is equivalent to 1.6606×10^{-24} g. *Use the Factor-Label Method* to calculate the mass (in grams) of a randomly-chosen sample of carbon containing 1 mole of carbon atoms. *Circle your answer*.

    b. Look at the Periodic Table. How does the number under the symbol for carbon compare to the number of grams you determined in Question 10a?

    c. What is the relationship between the units "atomic mass units" and "grams per mole"?

VIII. The Mole Concept – Start Time: __________ End Time: __________

Important!! For answers that involve a calculation you must show your work neatly using dimensional analysis with correct significant figures and units to receive full credit. No work, no credit. Report numerical answers to the correct number of significant figures. **CIRCLE ALL NUMERICAL RESPONSES**.

1. The atomic mass of Cl is 35.45 amu and that of Al is 26.98 amu. What are the masses in grams of…

   a.) 2.0 mol of Al atoms? *Circle your answer.*

   b.) 3.0 mol of Cl atoms? *Circle your answer.*

   c.) 3.0 mol chlorine molecules? *Circle your answer.*

2. a.) Why might the expression "1.00 mol of nitrogen" be confusing?

   b.) What change would remove any uncertainty?

   c.) For what other elements might a similar confusion exist? Why?

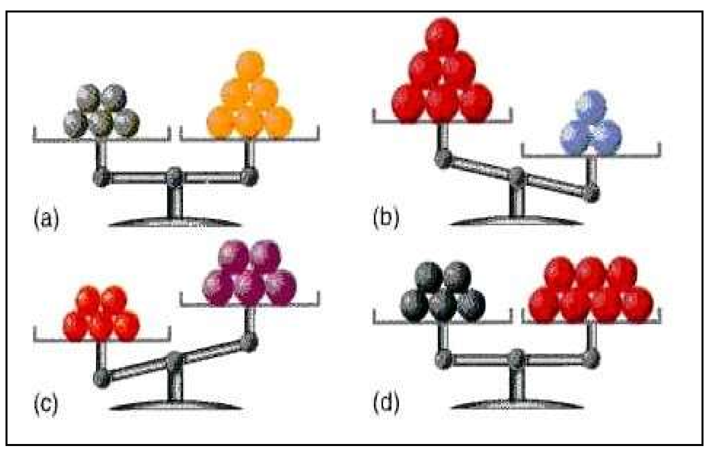
3. Each of the following balances weighs the indicated number of atoms of two elements. Which element, *left*, *right* or *neither*, has…

   a.) the higher molar mass on balance "a"?

   b.) more atoms per gram on balance "b"?

   c.) fewer atoms per gram on balance "c"?

   d.) more atoms per mole on balance "d"?



4. You need to calculate the number of P_4 molecules that can form from 2.5g of $Ca_3(PO_4)_2$. Explain how you would proceed, that is, write a solution "plan" without doing any calculations.

5. Calculate the ***molar mass*** of each of the following to *two decimal places*.

a.) Dinitrogen tetroxide *Circle your answer.*

b.) Calcium acetate *Circle your answer.*

6. Calculate each of the following quantities:

a.) The number of moles of chlorine atoms in 0.0425 g of $C_2H_4Cl_2$. *Circle your answer.*

b.) The total number of ions in 38.1 g calcium fluoride, CaF_2. *Circle your answer.*

c.) Mass in grams of 3.52 mol of chromium (III) sulfate decahydrate. *Circle your answer.*

d.) Mass in grams of 9.64×10^{24} molecules of dichlorine heptaoxide. *Circle your answer.*

e.) Number of moles and formula units in 56.2 g of lithium sulfate. *Circle your answer.*

f.) Number of lithium ions, sulfate ions, S atoms and O atoms in the mass in the previous problem (i.e. in 56.2 g of lithium sulfate) *Circle each answer.*

Li^+ ions:

S atoms:

Oxygen atoms:

7. Calculate the mass % of iodine, I, in strontium periodate. *Circle your answer.*

8. Oxygen is required for metabolic combustion of foods. Calculate the number of atoms in 38.0 g of oxygen gas, the amount absorbed into the blood from the lungs at rest in 15 minutes. *Circle your answer.*

9. Propane, C_3H_8, is widely used in liquid form as a fuel for barbecue grills and camp stoves. For 75.3 g of propane, calculate the following:
   a.) the moles of the compound in the sample *Circle your answer.*

   b.) the grams of carbon in the sample *Circle your answer.*

10. The effectiveness of a nitrogen fertilizer is determined mainly by the % N. Calculate the %N in each fertilizer and then rank them in terms of their effectiveness (i.e by their %N).
    a. potassium nitrate *Circle your answer.*

    b. ammonium nitrate *Circle your answer.*

    c. ammonium sulfate *Circle your answer.*

    d. urea, $CO(NH_2)_2$ *Circle your answer.*

Ranking by % N: ____

IX. Balancing Redox Reactions – Start Time: __________ End Time: __________

Oxidation-reduction or **Redox reactions** involve the transfer of one or more electrons from one chemical species to another. Redox reactions are involved in the corrosion of metals, the combustion of fuels, the generation of electricity from batteries and many biological processes including cellular respiration and photosynthesis. An understanding of redox chemistry is essential in the design of new kinds of batteries, increasing efficiency in fuel combustion, the prevention of corrosion, etc.

The **oxidation number** of an atom is the "apparent" charge the atom would have if each of its bonding electrons were assigned to the more electronegative atom in each bond. Oxidation numbers are useful in determining the substance **oxidized** (**L.E.O.** = **L**oss of **E**lectron(s) is **O**xidation) and the substance **reduced** (**G.E.R.** = **G**ain of **E**lectron(s) is **R**eduction). Hence, the substance that is oxidized loses electrons and therefore serves as a **reducing agent** since it provides electrons to another atom thereby causing that atom to be reduced. The species being reduced serves as the **oxidizing agent** because it removes electrons from another substance, thereby causing that substance to be oxidized.

<u>Rules to Assign Oxidation Numbers</u>

1. The oxidation number of any uncombined element is zero.
2. The oxidation number of a monatomic ion equal the charge on the ion.
3. The more electronegative element in a binary compound is assigned the number equal to the charge it would have if it were an ion.
4. Oxygen has an oxidation number of –2 unless it is combined with F, when it is usually +2, or it is in a peroxide, when it is –1.
5. The oxidation state of hydrogen in most of its compounds is +1 unless it combined with a metal, in which case it is –1.
6. In compounds, the elements of groups 1 and 2 as well as aluminum have oxidation number of +1, +2, and +3, respectively
7. The sum of the oxidation numbers of all atoms in a neutral compound is zero.
8. The sum of the oxidation number of all atoms in a polyatomic ion equals the charge of the ion.

Key Questions

1. The substances oxidized and reduced may not be obvious in the following redox reaction:

$$8\ H^{+}(aq) + Cr_2O_7^{2-}(aq) + 3\ SO_3^{2-}(aq) \rightarrow 2\ Cr^{3+}(aq) + 3\ SO_4^{2-}(aq) + 4\ H_2O(l)$$

   a. Write the oxidation numbers above each atomic symbol on the left and right sides of the reaction above.

   b. Does the oxidation number of any atom *increase*? If so, which one and what is the change?

   c. Does the oxidation number of any atom *decrease*? If so, which one and what is the change?

   d. Which species is being oxidized (oxidation number is increasing)? ____________________

   e. Which species is being reduced (oxidation number is decreasing)? ____________________

Exercises

2. Write the oxidation number above each symbol in the following compounds or ions.

   a. KBr b. BrF_3 c. $HBrO_3$ d. CBr_4 e. MnO_4^{1-} f. Mn_2O_3 g. $KMnO_4$

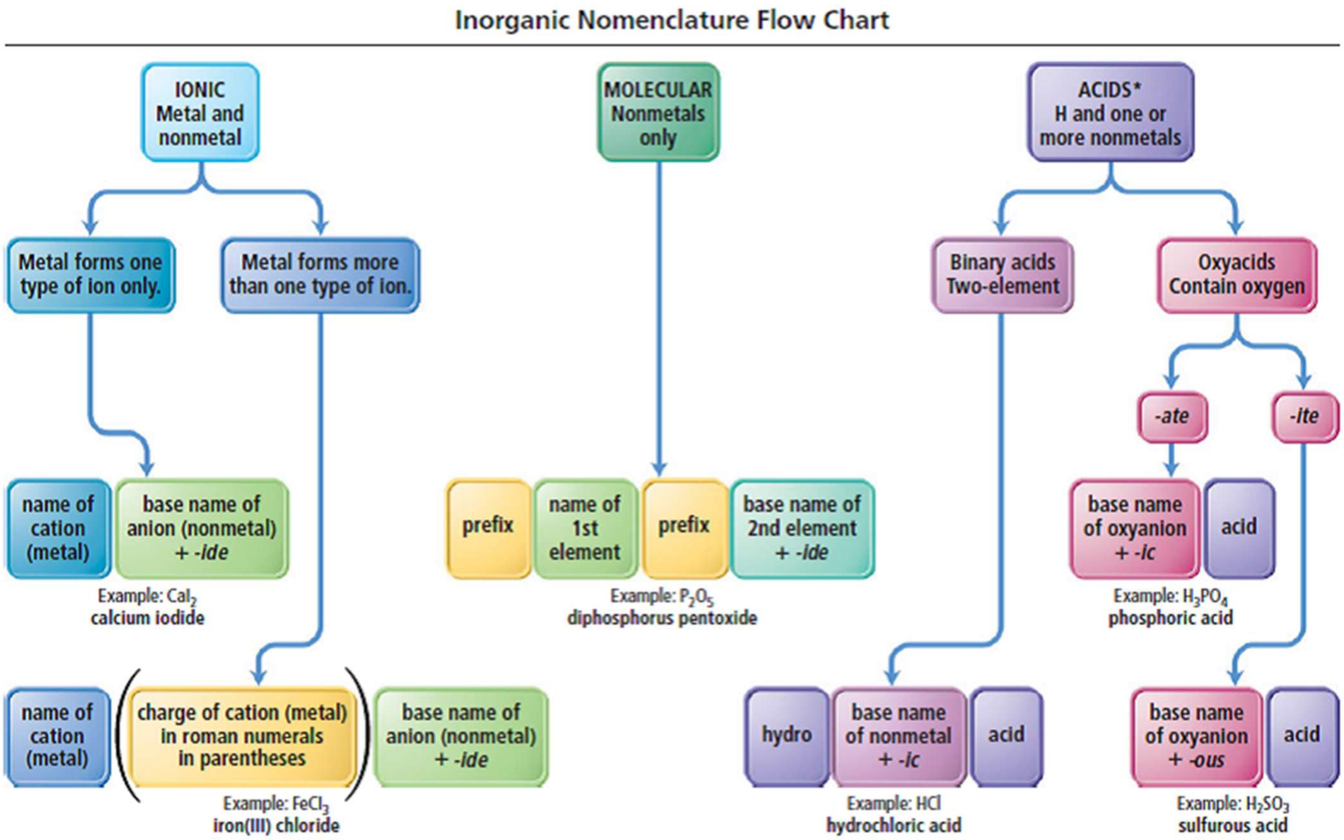
3. Assign oxidation numbers to each atom in the following reaction and then identify the oxidizing agent, reducing agent, the substance oxidized and the substance reduced.

$$NO_3^{1-}(aq) + 4\ Zn(s) + 7\ OH^{1-}(aq) + 6\ H_2O(l) \rightarrow 4\ Zn(OH)_4^{2-}(aq) + NH_3(aq)$$

Oxidizing agent: ________________ Reducing agent: ________________

Substance oxidized: _____________ Substance reduced: ______________

X. Chemical Nomenclature Flow Chart - Ungraded



▲ FIGURE 3.10 Inorganic Nomenclature Flow Chart The chart summarizes how to name inorganic compounds. Begin by determining if the compound to be named is ionic, molecular, or an acid. Then follow the flow chart for that category from top to bottom until you arrive at a name for the compound.

Case 1: Ionic compounds containing monatomic ions (i.e. ions that can only have one charge)

- Name of Compound = name of metal + name of non-metal w/ -ide suffix or name of polyatomic ion. *<u>No prefixes are used!</u>*
- e.g. NaF = sodium fluoride; Na_3PO_4 = sodium phosphate; $(NH_4)_3PO_4$ = ammonium phosphate

Case 2: Ionic compounds containing a metal that can form more than one ion

- Name of Compound = name of metal, followed by charge of metal in *Roman numerals in parentheses*, followed by name of non-metal w/ -ide suffix or name of polyatomic ion. *<u>No prefixes are used!</u>*
- e.g. $PbCl_2$ = lead (II) chloride; $Cu(NO_3)_2$ = copper (II) nitrate

Case 3: Binary molecular compounds

- Name of Compound = name of first element + name of second element with -ide suffix.
- Use prefixes (mono-, di-, tri-, tetra-, penta-, hexa-, hepta-, nona-, deca-) to indicate the number of atoms.
- The mono prefix is not used with the first element.
- e.g. CO = carbon monoxide; NO_2 = nitrogen dioxide; N_2O = dinitrogen monoxide; P_2O_5 = diphosphorous pentoxide

Case 4: Binary acid solutions (i.e. binary acids dissolved in water = binary acids in aqueous solution)

- Name of Compound = hydro + name of halogen w/ -ic suffix e.g. $HF_{(aq)}$ = hydrofluoric acid; $HCl_{(aq)}$ = hydrochloric acid
- Unless stated otherwise assume the formula of a binary acid is for the acid dissolved in water. e.g. assume $HCl = HCl_{(aq)}$

Naming Oxoacids (i.e. compound with the general formula H_xMO_y, where M = nonmetal)

The name of an oxoacid is based on the name of the polyatomic ion from which the acid is derived.

Case 5: -ate → -ic

If the name of the polyatomic ion ends in **"-ate**," the name of the corresponding acid ends in **"-ic** acid."

Polyatomic ion (-ate)	→	**Acid (-ic)**
sulf<u>ate</u> = SO_4^{2-}	→	H_2SO_4 = sulfur<u>ic</u> acid
Chlor<u>ate</u> = ClO_3^{1-}	→	$HClO_3$ = chlor<u>ic</u> acid

Case 6: -ite → -ous

If the name of the polyatomic ion ends in **"-ite**," the name of the corresponding acid ends in **"-ous** acid."

Polyatomic ion (-ite)	→	**Acid (-ous)**
sulf<u>ite</u> = SO_3^{2-}	→	H_2SO_3 = sulfur<u>ous</u> acid
Chlor<u>ite</u> = $ClO2^{1-}$	→	HClO2 = chlor<u>ous</u> acid

<u>Practice Exercises</u>

The following questions are for practice and will *<u>not</u>* be graded. See/email me if you want your answers checked.

1. Write the formula for the following polyatomic ions:
   a. Ammonium
   b. Hydroxide
   c. Nitrite
   d. Nitrate
   e. Acetate
   f. Cyanide
   g. Carbonate
   h. Sulfite
   i. Sulfate
   j. Phosphate

2. Write the formula fo the ionic compound that forms when the following ions combine:
   a. Na^+ and Cl^-
   b. Na^+ and O^{2-}
   c. Cu^{2+} and OH^-
   d. Ca^{2+} and CO_3^{2-}
   e. Sr^{2+} and PO_4^{3-}
   f. Al^{3+} and NO_3^-

3. Name the following ionic compounds.
   a. $MgCl_2$
   b. CaO
   c. Cu_3N_2
   d. AuF_3
   e. Ag_3PO_4
   f. NH_4Cl
   g. $NzSO_4$
   h. $Fe_2(SO_3)_3$
   i. $KC_2H_3O_2$
   j. NaI

4. Write the formula for the following ionic compounds.
   a. Lithium bromide
   b. Calcium carbonate
   c. Beryllium nitride
   d. Potassium nitrate
   e. Copper (II) sulfite
   f. Gold (I) phosphate
   g. Cobalt (III) oxide
   h. Calcium acetate
   i. Iron (III) cyanide
   j. Aluminum hydroxide

5. Name the following ionic compounds:
   a. $Fe(NO_3)_3$
   b. $MgBr_2$
   c. Au_2S
   d. Na_3PO_4
   e. K_4C
   f. $ZnCl_2$
   g. $Co_2(SO_4)_3$
   h. Al_2S_3
   i. CaS
   j. NiO_2

6. Write the formula for the following ionic compounds
   a. Iron (II) acetate
   b. Copper (I) oxide
   c. Gold (III) nitride
   d. Calcium phosphate
   e. Potassium sulfate
   f. Ruthenium (II) nitrate
   g. Sodium chloride
   h. Lithium sulfate
   i. Beryllium sulfite
   j. Aluminum carbonate

7. Write the name of the following covalent compounds:
   a. N_2O
   b. CO_2
   c. CO
   d. NO
   e. P_2S
   f. NBr_3
   g. IBr
   h. CF_4
   i. P_2F_6
   j. C_2S_4

8. Write the formula for the following covalent compounds
   a. Nitrogen trisulfide
   b. Oxygen difluoride
   c. Diphosphorous pentoxide
   d. Sulfur dichloride
   e. Nitrogen triiodide
   f. Disulfur tetrachloride
   g. Dicarbon hexabromide
   h. Tricarbon octafluoride
   i. Dihydrogen monoxide
   j. Tetrahydrogen monocarbide

9. Write the name for the following acids
   a. N_sSO_4
   b. HCl
   c. HBr
   d. HNO_3
   e. HF
   f. H_3PO_4
   g. H_2SO_3
   h. HNO_2
   i. HCN

10. Write the formula for the following acids
    a. Nitrous acid
    b. Hydrochloric acid
    c. Hydrobromic acid
    d. Sulfuric acid
    e. Sulfurous acid
    f. Acetic acid
    g. Carbonic acid
    h. Perchloric acid

11. Mixed Bag – name the following compounds
    a. MgI_2
    b. $HClO_4$
    c. C_3Cl_8
    d. Co_2S_3
    e. $AgNO_3$
    f. $CuSO_4$
    g. HBr
    h. $NaCl$
    i. ZnS

12. Mixed Bag – give the formula for the following compounds
    a. Gold (III) sulfide
    b. Hydrofluoric acid
    c. Aluminum oxide
    d. Magnesium sulfate
    e. Nitric acid
    f. Iron (III) acetate
    g. Silver nitrate
    h. Potassium phosphate
    i. Dicarbon hexafluoride

13. Hydrates – name the following compounds (normal nomenclature rules and then add a covalent prefix to the word hydrate i.e. $AlCl_3 \bullet 6H_2O$ is aluminum chloride hexahydrate)
    a. $CaSO_4 \bullet 2H_2O$
    b. $CuSO_4 \bullet 5H_2O$
    c. $MgSO_4 \bullet 7H_2O$
    d. $Ma(OH)_2 \bullet 8H_2O$

14. Hydrates – give the formula for the following compounds (put a ● between the compound and the water molecules)
    a. Copper (III) nitrate trihydrate
    b. Barium chloride dihydrate
    c. Cobalt (II) nitrate hexahydrate
    d. Cobalt (II) sulfate heptahydrate
    e. Iron (III) sulfate pentahydrate

XI. Elements and Symbols to Memorize

It is expected that you know the symbols and names (spelling too!) for the following elements from your study in a previous chemistry class. <u>You will be tested on these symbols on the first AP Chem exam</u>.

ELEMENT	**SYMBOL**
Aluminum	Al
Antimony	Sb
Argon	Ar
Arsenic	As
Barium	Ba
Beryllium	Be
Bismuth	Bi
Boron	B
Bromine	Br
Cadmium	Cd
Calcium	Ca
Carbon	C
Chlorine	Cl

ELEMENT	**SYMBOL**
Chromium	Cr
Cobalt	Co
Copper	Cu
Fluorine	F
Gold	Au
Helium	He
Hydrogen	H
Iodine	I
Iron	Fe
Krypton	Kr
Lead	Pb
Lithium	Li
Magnesium	Mg

ELEMENT	**SYMBOL**
Manganese	Mn
Mercury	Hg
Molybdenum	Mo
Neon	Ne
Nickel	Ni
Nitrogen	N
Oxygen	O
Phosphorous	P
Platinum	Pt
Plutonium	Pu
Potassium	K
Radium	Ra
Radon	Rn

ELEMENT	**SYMBOL**
Selenium	Se
Silicon	Si
Silver	Ag
Sodium	Na
Strontium	Sr
Sulfur	S
Tin	Sn
Titanium	Ti
Tungsten	W
Uranium	U
Vanadium	V
Xenon	Xe
Zinc	Zn

XII. Ion Sheet to Memorize (Group IA are 1+, Group IIA are 2+, Group VIA are 2-, Group VIIA are 1-)
You should also know polyatomic ion rules (e.g. XO_4^-=per__ate, XO_3^-= __ate, XO_2^-= __ite, XO^-=hypo__ite)

Cations

1+	
Ammonium	NH_4^+
Copper (I)	Cu^+
Hydrogen	H^+
Hydronium	H_3O^+
Lithium	Li^+
Mercury (I)	Hg_2^{+2}
Potassium	K^+
Silver	Ag^+
Sodium	Na^+

2+	
Barium	Ba^{2+}
Calcium	Ca^{2+}
Cadmium	Cd^{2+}
Chromium (II)	Cr^{2+}
Cobalt (II)	Co^{2+}
Copper (II)	Cu^{2+}
Iron (II)	Fe^{2+}
Lead (II)	Pb^{2+}
Magnesium	Mg^{2+}
Manganese (II)	Mn^{2+}
Mercury (II)	Hg^{2+}
Nickel (II)	Ni^{2+}
Tin (II)	Sn^{2+}
Zinc	Zn^{2+}

3+	
Antimony (III)	Sb^{3+}
Arsenic (III)	As^{3+}
Bismuth (III)	Bi^{3+}
Chromium III)	Cr^{3+}
Cobalt (III)	Co^{3+}
Iron (III)	Fe^{3+}
Nickel (III)	Ni^{3+}

4+	
Lead (IV)	Pb^{4+}
Manganese (IV)	Mn^{4+}
Tin (IV)	Sn^{4+}

5+	
Antimony (V)	Sb^{5+}
Bismuth (V)	Bi^{5+}

Anions

1-	
Acetate	CH_3COO^-
Bicarbonate	HCO_3^-
Binoxalate	$HC_2O_4^-$
Bisulfate	HSO_4^-
Bisulfide	HS^-
Bisulfite	HSO_3^-
Bromide	Br^-
Chlorate	ClO_3^-
Chloride	Cl^-
Chlorite	ClO_2^-
Cyanide	CN^-
Dihydrogen phosphate	$H_2PO_4^-$
Fluoride	F^-
Hydroxide	OH^-
Hypochlorite	ClO^-
Iodide	I^-
Nitrate	NO_3^-
Nitrite	NO_2^-
Perchlorate	ClO_4^-
Permanganate	MnO_4^-
Thiocyanate	SCN^-

2-	
Carbonate	CO_3^{2-}
Chromate	CrO_4^{2-}
Dichromate	$Cr_2O_7^{2-}$
Monohydrogen phosphate	HPO_4^{2-}
Oxalate	$C_2O_4^{2-}$
Oxide	O^{2-}
Peroxide	O_2^{2-}
Sulfate	SO_4^{2-}
Sulfide	S^{2-}
Sulfite	SO_3^{2-}

3-	
Arsenate	AsO_4^{3-}
Borate	BO_3^{3-}
Ferricyanide	$Fe(CN)_6^{3-}$
Nitride	N^{3-}
Phosphate	PO_4^{3-}
Phosphide	P^{3-}

XIII. General Chemistry Review – Graded on Completion

The following questions are just covering topics that should be easily recalled. Please try this section without looking at the previous parts of this packet or any other notes that you may have. There is a periodic table and a list of equations at the end of this packet; you can use that if you need to. If, after you have finished, you would like to go back and make sure you got the correct answers, feel free to do that.

_____ 1. Consider the following selected postulates of Dalton's atomic theory:
I. Each element is composed of extremely small particles called atoms.
II. Atoms are indivisible
IIII. Atoms of a given element are identical.
IV. Atoms of different elements are different and have different properties.
Which of the postulates is/are no longer considered valid?

a. I and II
b. II only
c. II and III
d. III only
e. III and IV

_____ 2. There are _____ electrons, _____ protons, and _____ neutrons in an atom of ${}^{132}_{54}X$.

a. 132, 132, 54
b. 54, 54, 132
c. 78, 78, 54
d. 54, 54, 78
e. 78, 78, 132

_____ 3. Which pair of atoms constitutes a pair of isotopes of the same element?

a. ${}^{14}_{6}X$ ${}^{14}_{7}X$
b. ${}^{14}_{6}X$ ${}^{12}_{6}X$
c. ${}^{17}_{9}X$ ${}^{17}_{8}X$
d. ${}^{19}_{10}X$ ${}^{19}_{9}X$
e. ${}^{20}_{10}X$ ${}^{21}_{11}X$

_____ 4. The element X has three naturally occurring isotopes. The masses (amu) and % abundance of the isotopes are given in the table below. The average atomic mass of the element is _____ amu.

Isotope	Abundance	Mass
${}^{221}X$	74.22	220.9
${}^{220}X$	12.78	220.0
${}^{218}X$	13.00	218.1

a. 219.7
b. 220.4
c. 220.42
d. 218.5
e. 221.0

_____ 5. Silver has two naturally occurring isotopes with the following atomic masses: ${}^{107}Ag$ (106.90509 amu) and ${}^{109}Ag$ (108.9047). The average atomic mass of silver is 107.8682 amu. The fractional abundance of the lighter of the two isotopes is _____.

a. 0.24221
b. 0.48168
c. 0.51835
d. 0.75783
e. 0.90474

_____ 6. Which of the following conclusions can be drawn from J.J. Thomson's plum pudding model?
- a. Atoms contain electrons.
- b. Practically all the mass of an atom is contained in its nucleus.
- c. Atoms contain protons, neutrons, and electrons.
- d. Atoms have a positively charged nucleus surrounded by an electron cloud.
- e. No two electrons in one atom can have the same four quantum numbers.

Questions 7-9 use the following answer choices:
- a. Wave nature of matter
- b. Spectral lines
- c. Quantum numbers
- d. Bohr model

_____ 7. Explained spectral lines by postulating that electrons were only able to exist in discrete orbits of differing energies around the atom.

_____ 8. Describes the location of an electron series of possible quantum states that are allowed, some of which are favored energetically for certain electrons over others.

_____ 9. Caused by electrons emitting energy as they transitioned from one specific orbit to another.

_____ 10. Which of the following is a correct interpretation of the results of Rutherford's experiments?
- a. Atoms have equal numbers of positive and negative charges.
- b. Electrons in atoms are arranged in shells.
- c. Neutrons are at the center of an atom.
- d. Neutrons and protons in atoms have nearly equal mass.
- e. The positive charge of an atom is concentrated in a small region.

_____ 11. The energy of a photon of light is ____ proportional to its frequency and ____ proportional to its wavelength.
- a. Directly, directly
- b. Inversely, inversely
- c. Inversely, directly
- d. Directly, inversely
- e. Inversely, not

_____ 12. What is the wavelength of light that has a frequency of 3.22×10^{14} s^{-1}?
- a. 932 nm
- b. 649 nm
- c. 9.66×10^{22} nm
- d. 9.32×10^{-7} nm
- e. 1.07×10^{6} nm

_____ 13. The energy of a photon that has a wavelength of 9.0 m is ____ J.
- a. 2.2×10^{-26}
- b. 4.5×10^{25}
- c. 6.0×10^{-23}
- d. 2.7×10^{9}
- e. 4.5×10^{-25}

_____ 14. Which pair of elements would you expect to exhibit the greatest similarity in their physical and chemical properties?

a. O, S
b. C, N
c. K, Ca
d. H, He
e. Si, P

_____ 15. An element that appears in the lower left corner of the periodic table is ____.

a. Either a metal or a metalloid
b. Definitely a medal
c. Either a metalloid or a non-metal
d. Definitely a non-metal
e. Definitely a metalloid

_____ 16. The elements in groups 1, 17, and 18 are called ____, respectively.

a. Alkaline earth metals, halogens, and chalcogens
b. Alkali metals, chalcogens, and halogens
c. Alkali metals, halogens, and noble gases
d. Alkaline earth metals, transition metals, and halogens
e. Halogens, transition metals

_____ 17. Which one of the following is most likely to lose electrons when forming an ion?

a. F
b. P
c. Rh
d. S
e. N

_____ 18. When a metal and a nonmetal react, the ____ tends to lose electrons and the ____ tends to gain electrons.

a. Metal, metal
b. Nonmetal, nonmetal
c. Metal, nonmetal
d. Nonmetal, metal
e. None of the above (these elements share electrons)

_____ 19. ____ typically form ions with a 2+ charge.

a. Alkaline earth metals
b. Halogens
c. Chalcogens
d. Alkali metals
e. Transition metals

_____ 20. What is the ratio of hydrogen to oxygen atoms in the mineral cacoxenite, $Fe_4(PO_4)_3(OH)_3 \bullet 12H_2O$?

a. 27:19
b. 15:27
c. 15:24
d. 1:1
e. 27:25

_____ 21. Which of the following compounds would you expect to be ionic?

a. SF_6
b. H_2O
c. H_2O_2
d. NH_3
e. CaO

_____ 22. The charge on the iron ion in the salt Fe_2O_3 is _____.

a. +1
b. +2
c. +3
d. -5
e. -6

_____ 23. What is the formula of the compound strontium nitride?

a. SrN
b. Sr_3N_2
c. Sr_2N_3
d. SrN_2
e. SrN_3

_____ 24. Which of the following is the nitride ion?

a. Na^+
b. NO_3^-
c. NO_2^-
d. NH_4^+
e. N^{3-}

_____ 25. What is the correct formula for chromium (III) oxide?

a. Cr_2O_3
b. CrO_3
c. Cr_3O_2
d. Cr_3O
e. Cr_2O_4

_____ 26. What is the correct formula for ammonium sulfide?

a. NH_4SO_3
b. $(NH_4)_2SO_3$
c. $(NH_4)_2S$
d. NH_3S
e. N_2S_3

_____ 27. Which formula is not correctly paired with its compound name?

a. $FeSO_4$ iron (II) sulfate
b. $Fe_2(SO_3)_3$ iron (III) sulfite
c. FeS iron (II) sulfide
d. $FeSO_3$ iron (II) sulfite
e. $Fe_2(SO_4)_3$ iron (III) sulfide

_____ 28. The formula for the compound formed between aluminum ions and phosphate ions is _____.

a. $Al_3(PO_4)_3$
b. $AlPO_4$
c. $Al(PO_4)_3$
d. $Al_2(PO_4)_3$
e. AlP

_____ 29. Barium reacts with a polyatomic ion to form a compound with the general formula $Ba_3(X)_2$. What would be the most likely formula for the compound formed between sodium and the polyatomic ion, X?

a. NaX
b. Na_2X
c. $Na_2(X)_2$
d. Na_3X
e. $Na_3(X)_2$

_____ 30. Which one of the following is true concerning the makeup of the simplest unit of magnesium chloride?
- a. 1 magnesium atom and 1 diatomic chlorine molecule
- b. 1 magnesium chloride molecule
- c. 1 magnesium atom and 2 chlorine atoms
- d. 1 positive ion and 2 negative ions
- e. 2 positive ions and 1 negative ion

_____ 31. The molar mass of ammonium phosphate is
- a. 116.03 g/mol
- b. 121.07 g/mol
- c. 149.09 g/mol
- d. 155.42 g/mol
- e. 242.01 g/mol

_____ 32. One mole of ____ contains the largest number of atoms.
- a. S_8
- b. $C_{10}H_8$
- c. $Al_2(SO_4)_3$
- d. Na_3PO_4
- e. Cl_2

_____ 33. A 30.5-gram sample of glucose ($C_6H_{12}O_6$) contains ____ mol of glucose.
- a. 0.424
- b. 0.169
- c. 5.90
- d. 2.36
- e. 0.136

_____ 34. How many moles of oxygen are in 1.08 moles of calcium nitrate?
- a. 7.55 moles
- b. 1.43 moles
- c. 6.48 moles
- d. 33.8 moles
- e. 1.16×10^{23} moles

_____ 35. How many grams of calcium nitrate contains 24 grams of oxygen atoms?
- a. 164 grams
- b. 96 grams
- c. 62 grams
- d. 50. grams
- e. 41 grams

_____ 36. How many molecules of methane, CH_4, are in 48.2 g of methane?
- a. 5.00×10^{24} molecules
- b. 3.00 molecules
- c. 2.90×10^{25} molecules
- d. 1.81×10^{24} molecules
- e. 4.00 molecules

_____ 37. What number of moles of diatomic oxygen is needed to produce 14.2 grams of tetraphosphorous decoxide from singlet phosphorous in a synthesis reaction?
- a. 0.0500 moles
- b. 0.0625 moles
- c. 0.125 moles
- d. 0.250 moles
- e. 0.500 moles

_____ 38. Which of the following gas samples contains the greatest mass of gas molecules (assume ideal conditions)?

a. 1.0 L of He at STP
b. 1.0 L of Xe at STP
c. 1.0 L of F_2 at STP
d. 1.0 L of H_2 at STP
e. All three are the same.

_____ 39. The molarity of an aqueous solution containing 75.3 g of glucose ($C_6H_{12}O_6$) in 35.5 mL of solution is ____.

a. 1.85 M
b. 2.12 M
c. 0.197 M
d. 3.52 M
e. 11.8 M

_____ 40. How many mililiters of 0.123 M NaOH solution contain 25.0 g of NaOH (molar mass = 40.00 g/mol)?

a. 50.80 mL
b. 625 mL
c. 5080 mL
d. 7.69 mL
e. 12.7 mL

_____ 41. What is the total concentration of ions in 0.0360 M solution of sodium carbonate?

a. 0.0900 M
b. 0.108 M
c. 0.0120 M
d. 0.0720 M
e. 0.0360 M

42. Radioactive americium-241 is used in household smoke detectors and in bone mineral analysis.

a. Give the number of electrons, protons, and neutrons in an atom of americium-241.

b. Write the proper nuclide symbol.

43. What characteristics do atoms of carbon-12, carbon-13, and carbon-14 have in common? In what ways are they different?

44. Identify the isotope that has atoms with

a. 117 neutrons, 77 protons, and 77 electrons

b. 30 neutrons, 28 protons, and 28 electrons

45. How did the discovery of isotopes conflict with Dalton's atomic theory?

46. The average mass of any large number of atoms of a given element is always the same for a given element. Explain.

47. Naturally occurring boron is 19.9% B-10 (10.01294 u) and 80.1% B-11 (11.0093 u). Calculate the average atomic mass.

48. Uranium has an atomic mass equal to 238.0289 amu. It consists of two isotopes: uranium-235 (2335.044 u) and uranium-238 (238.051 u). Calculate the % abundance of the uranium-235 isotope.

49. How did Rutherford interpret the deflection of α-particles in his gold foil experiment? Did these findings support or disprove the plum pudding model of the atom? Explain.

50. How did Bohr's model of the atom explain the existence of line spectra? How are spectral lines produced?

51. What is the energy of a photon that has a wavelength of 8.33×10^{-6} m?

52. What is the wavelength of a photon that has an energy of 5.25×10^{-19} J?

53. How does the quantum model describe the location of an electron?

54. What accounts for similarities of chemical properties for elements in the same group/family?

55. Provide the group names for the elements in Group 1, 2, 17, and 18. Provide an example of an element in each of the above groups.

56. Identify the following elements:
    a. A halogen in the 3rd period
    b. A metalloid
    c. An atom in the 4th period that forms a stable ion with a +1 charge

57. Locate the following elements on the periodic table and indicate which orbital type is occupied by its valence electrons
    a. Lithium
    b. Silicon
    c. Copper

58. What ions are the following elements likely to form?
    a. Oxygen
    b. Sodium
    c. Bromine

59. A main group element in Period 4 forms the molecule compound H_2E and the ionic compound Na_2E.
    a. To which group does the element belong?
    b. Write the name and symbol of the element.

60. Write the formulas for the following binary ionic compounds.
    a. Sodium sulfide
    b. Cobalt (II) chloride
    c. Lithium nitride
    d. Tin (IV) oxide

61. Write the formulas for the following polyatomic ionic compounds
    a. Barium nitrate
    b. Calcium phosphite
    c. Iron (II) chromate
    d. Potassium permanganate

62. How many total ions are present in the following ionic compounds?
    a. Sodium acetate
    b. Aluminum nitrate
    c. Copper (II) chloride

63. The most common charge associate with silver in its compounds is +1. Indicate the formulas you would expect for the ionic compounds formed between silver and the following elements.
    a. Iodine
    b. Sulfur
    c. Phosphorous

64. Answer the following questions for a 3.50 g sample of glucose, $C_6H_{12}O_6$.
    a. What is the molar mass of this compound?
    b. How many moles are in the sample?
    c. How many hydrogen atoms are present in the sample?

65. Calculate the number of moles in 1.75 grams of sodium carbonate.
    a. How many formula units are present?
    b. How many sodium ions are present?

66. Boron has two isotopes B-10 (19.9%) and B-11 (80.1%).
    a. How many atoms of B would be present in a 50-gram sample of pure boron?

    b. How many of these atoms are B-10?

67. A sample of $C_{12}H_{22}O_{11}$ contains 0.4662 moles of carbon atoms. How many moles of hydrogen atoms are in the sample?

68. Without doing any detailed calculations, rank the following samples in order of increasing number of atoms:
    a. 3.0×10^{23} molecules H_2O_2
    b. 2 mol CH_4
    c. 32 g O_2

69. One component of smog is nitrogen monoxide, NO. A car produces about 8.0 g of this gas per day. What is the volume at STP?

70. A reaction produces 100 grams of water.
    a. How many grams of H_2 must have reacted to produce this amount of water if 1 mol of H_2O is produced for every 1 mol of H_2 that reacts?

    b. Assuming H_2 gas was the entire source of H and all of it was converted to water, how many L of H_2 gas reacted, assuming the reaction was carried out at STP conditions?

    c. How many molecules of H_2 reacted, assuming the reaction was carried out at STP conditions?

71. Calculate the following quantities for 343 mL of a 1.27 M Na_2SO_4 solution:
    a. Moles of sodium sulfate

    b. Grams of sodium sulfate required to prepare the above solution.

    c. Moles of Na^+

72. What is the molarity of a 750.-mL solution containing 50.0 g of potassium chloride?

73. How many moles of hydroxide ions are present in 300. mL of a 2.50 M $Ca(OH)_2$ solution?

    a. What is the molarity of hydroxide ions?

PERIODIC TABLE OF THE ELEMENTS

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1 H 1.008																	2 He 4.00
3 Li 6.94	4 Be 9.01											5 B 10.81	6 C 12.01	7 N 14.01	8 O 16.00	9 F 19.00	10 Ne 20.18
11 Na 22.99	12 Mg 24.30											13 Al 26.98	14 Si 28.09	15 P 30.97	16 S 32.06	17 Cl 35.45	18 Ar 39.95
19 K 39.10	20 Ca 40.08	21 Sc 44.96	22 Ti 47.87	23 V 50.94	24 Cr 52.00	25 Mn 54.94	26 Fe 55.85	27 Co 58.93	28 Ni 58.69	29 Cu 63.55	30 Zn 65.38	31 Ga 69.72	32 Ge 72.63	33 As 74.92	34 Se 78.97	35 Br 79.90	36 Kr 83.80
37 Rb 85.47	38 Sr 87.62	39 Y 88.91	40 Zr 91.22	41 Nb 92.91	42 Mo 95.95	43 Tc	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.76	52 Te 127.60	53 I 126.90	54 Xe 131.29
55 Cs 132.91	56 Ba 137.33	57-71 *	72 Hf 178.49	73 Ta 180.95	74 W 183.84	75 Re 186.21	76 Os 190.23	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po	85 At	86 Rn
87 Fr	88 Ra	89-103 †	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Mc	116 Lv	117 Ts	118 Og

*Lanthanoids	57 La 138.91	58 Ce 140.12	59 Pr 140.91	60 Nd 144.24	61 Pm	62 Sm 150.36	63 Eu 151.97	64 Gd 157.25	65 Tb 158.93	66 Dy 162.50	67 Ho 164.93	68 Er 167.26	69 Tm 168.93	70 Yb 173.05	71 Lu 174.97
†Actinoids	89 Ac	90 Th 232.04	91 Pa 231.04	92 U 238.03	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

AP® CHEMISTRY EQUATIONS AND CONSTANTS

Throughout the exam the following symbols have the definitions specified unless otherwise noted.

L, mL	= liter(s), milliliter(s)	mm Hg	= millimeters of mercury
g	= gram(s)	J, kJ	= joule(s), kilojoule(s)
nm	= nanometer(s)	V	= volt(s)
atm	= atmosphere(s)	mol	= mole(s)

ATOMIC STRUCTURE

$E = h\nu$

$c = \lambda\nu$

E = energy
ν = frequency
λ = wavelength

Planck's constant, $h = 6.626 \times 10^{-34}$ J s
Speed of light, $c = 2.998 \times 10^{8}$ m s^{-1}
Avogadro's number = 6.022×10^{23} mol^{-1}
Electron charge, $e = -1.602 \times 10^{-19}$ coulomb

EQUILIBRIUM

$$K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}, \text{ where } a\,A + b\,B \rightleftarrows c\,C + d\,D$$

$$K_p = \frac{(P_C)^c(P_D)^d}{(P_A)^a(P_B)^b}$$

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14} \text{ at } 25^\circ C$$
$$= K_a \times K_b$$

$$pH = -\log[H^+],\ pOH = -\log[OH^-]$$

$$14 = pH + pOH$$

$$pH = pK_a + \log\frac{[A^-]}{[HA]}$$

$$pK_a = -\log K_a,\ pK_b = -\log K_b$$

Equilibrium Constants

K_c (molar concentrations)
K_p (gas pressures)
K_a (weak acid)
K_b (weak base)
K_w (water)

KINETICS

$$[A]_t - [A]_0 = -kt$$

$$\ln[A]_t - \ln[A]_0 = -kt$$

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$t_{1/2} = \frac{0.693}{k}$$

k = rate constant
t = time
$t_{1/2}$ = half-life

GASES, LIQUIDS, AND SOLUTIONS

$$PV = nRT$$

$$P_A = P_{total} \times X_A, \text{ where } X_A = \frac{\text{moles A}}{\text{total moles}}$$

$$P_{total} = P_A + P_B + P_C + \ldots$$

$$n = \frac{m}{\boldsymbol{M}}$$

$$\text{K} = {}^\circ\text{C} + 273$$

$$D = \frac{m}{V}$$

$$KE_{\text{molecule}} = \frac{1}{2}mv^2$$

Molarity, M = moles of solute per liter of solution

$$A = \varepsilon bc$$

P = pressure
V = volume
T = temperature
n = number of moles
m = mass
$\boldsymbol{M}$ = molar mass
D = density
KE = kinetic energy
v = velocity
A = absorbance
ε = molar absorptivity
b = path length
c = concentration

Gas constant, R = 8.314 J mol^{-1} K^{-1}
= 0.08206 L atm mol^{-1} K^{-1}
= 62.36 L torr mol^{-1} K^{-1}

1 atm = 760 mm Hg = 760 torr

STP = 273.15 K and 1.0 atm

Ideal gas at STP = 22.4 L mol^{-1}

THERMODYNAMICS/ELECTROCHEMISTRY

$$q = mc\Delta T$$

$$\Delta S^\circ = \sum S^\circ \text{ products} - \sum S^\circ \text{ reactants}$$

$$\Delta H^\circ = \sum \Delta H_f^\circ \text{ products} - \sum \Delta H_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \sum \Delta G_f^\circ \text{ products} - \sum \Delta G_f^\circ \text{ reactants}$$

$$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$$
$$= -RT\ln K$$
$$= -nFE^\circ$$

$$I = \frac{q}{t}$$

$$E_{cell} = E_{cell}^\circ - \frac{RT}{nF}\ln Q$$

q = heat
m = mass
c = specific heat capacity
T = temperature
S° = standard entropy
H° = standard enthalpy
G° = standard Gibbs free energy
n = number of moles
E° = standard reduction potential
I = current (amperes)
q = charge (coulombs)
t = time (seconds)
Q = reaction quotient

Faraday's constant, F = 96,485 coulombs per mole of electrons

$$1 \text{ volt} = \frac{1 \text{ joule}}{1 \text{ coulomb}}$$

XVI. End of Packet Survey

Please fill in the following so that I know what I need to go over the first week.

For this section, please rate from 1-10 your understanding of the topic (1 being the least, 10 being the most)

TOPIC	LEAST UNDERSTANDING							MOST UNDERSTANDING		
III	1	2	3	4	5	6	7	8	9	10
IV	1	2	3	4	5	6	7	8	9	10
V	1	2	3	4	5	6	7	8	9	10
VI	1	2	3	4	5	6	7	8	9	10
VII	1	2	3	4	5	6	7	8	9	10
VIII	1	2	3	4	5	6	7	8	9	10
IX	1	2	3	4	5	6	7	8	9	10

For this section, please tell me how long you spent doing each section.

III	__________ hours	VII	__________ hours
IV	__________ hours	VIII	__________ hours
V	__________ hours	IX	__________ hours
VI	__________ hours		

Do you have access to a non-Chromebook computer at home? ________ Yes _________ No

Please include any other thoughts on the back of this sheet.