Student Name: ______________________

Chemistry
Summer Work Packet

All students taking chemistry course are required to complete a review packet prior to the start of the course. This packet is designed to help the student review material that was learned in prerequisite science and math classes. A pretest will be administered the first day of class to assess the students' knowledge of the science concepts covered in the packet.

I will like to have a personal meeting with each of the students parents at the beginning of the year to discuss:

1. The strength and weakness of each student
2. How can we help the student as a team and improve their weakness and take advantages of their strengths.
3. The performance of the student in the summer work packet due on the first day of class.

I have read and understand the information written above.

Student signature: ______________________________________ Date ______________

Parent/guardian signature: _______________________________ Date ______________

I attest that all of the work contained in this packet is my own. *This does not mean that you can't work together on this, I HIGHLY SUGGEST IT, however this means that you did not just copy someone's work.*

Student signature: ______________________________________ Date ______________

PARENTS: Please take some time and look through this packet as well. Pay particular attention to the "Show your work" and "Desired qualities of a Saint Mary's student" sections!

In order to prepare for your semester in Chemistry, a packet of information and work has been assigned over the summer. This packet is split in to eight parts, each of which focuses on a different aspect. There is additional packet of work/answer sheets, which will be collected on the first day of the semester.

This packet should be completed during the summer. There will be an accompanying test on the concepts involved in this packet shortly after the year begins. Students will have an opportunity to ask questions on the packet, however the teacher will NOT directly address every single topic. Keep in mind that this is the information that you should come into the course knowing. Almost all of the information presented should not be new to you.

It is expected that this packet will be completed in a timely manner, and not at the last minute. It is recommended that this packet should be looked through for an understanding of what will be expected of you when you begin Chemistry. If the material seems overwhelming difficult do not give up, become a researcher and continue investigating the topic. Information is provided in the packet, but information should also be sought from other resources, books or reliable websites on the internet. Many colleges, educational institutions, and I-tunes, youtube has relevant information on the topics discussed in this packet. Please use them. Internet access is not required to complete this packet, but it will help.

If there are any questions or if you need clarifications over the summer, you may e-mail Mr. Lopez at alopez@stmaryhs.org, but be forewarned that over the summer email is not always checked immediately. Attempt to seek information early, rather than waiting until the end.

Regards and God bless you,

Mr. Lopez

CONTENTS

SECTION 1 – SHOW YOUR WORK HANDOUT

SHOW YOUR WORK

What does SHOW YOUR WORK even mean? You see it everywhere. It means different things to different people. But when in Chemistry, SHOW YOUR WORK means something very specific.

When showing work, you're describing a narrative, giving a step by step recipe for solving a problem. Even if you know how to solve the problem in your head, SHOW YOUR WORK means that you need to know how to express that know-how onto paper. It's a way of explaining your thought processes- even the ones you don't realize that you have. It is a systematic way of describing your work. And on top of that, if a person grading your work does not understand what it is you're trying to do, they will give up and you won't get to take part in any of that sweet partial credit everyone always talks about. Often times, poorly shown work will even result in a loss of credit, all because
SHOW YOUR WORK is a very specific statement.

I'll use an example, and you may not understand the problem, but the step by step process is how to solve it.

How many moles of Sodium are in a 120.0g sample of Sodium?

Step 1: Identify Variables and Constants
To perform this calculation, write out what you're given and identify what dimension the value measures. Include units and give the number as written (to keep significant figures).

Mass = 120.0 g

Also, other information is provided. Though you will learn about it this year, with the periodic table, knowing that the substance is sodium will give you that the Molar Mass of Sodium is 22.99 g/mol. Even though this isn't a variable, it is a constant (or tabulated value) so you should list it as well:

Mass = 120.0 g
Molar Mass = 23.0 g/mol (we always round our molar masses to one decimal)

Last, identify what it is you're trying to find. You can do this by writing the dimension you're looking for and signal it's the missing one with a "?".

Mass = 120.0 g
Molar Mass = 23.0 g/mol
n (moles) = ??

So now you've listed out your 'givens,' you can either use this to identify what equation to use, or you can simply state the equation. Write the equation out that you're going to use.

Mass = 120.0 g	**Molar Mass = mass/moles**
Molar Mass = 23.0 g/mol	
n (moles) = ??	

In this case, we're using the Molar Mass equation where Molar Mass equals mass over moles.

Now, beneath the used equation, rearrange the equation to solve for the unit you're trying to find. Do this BEFORE you input your numbers in, so that you can see the proper rearrangement of the equation before it becomes a mess:

Mass = 120.0 g	Molar Mass = mass/moles
Molar Mass = 23.0 g/mol	**Moles = mass/molar mass**
n (moles) = ??	

This requires algebra, but it's easier to do algebra with letters than with numbers and units.

Once you have the variables declared and the equation solved for the variable you want to find, plug the numbers in:

Mass = 120.0 g	Molar Mass = mass/moles
Molar Mass = 23.0 g/mol	Moles = mass/molar mass
n (moles) = ??	**Moles =** $\mathbf{\frac{120.0\ g}{23.0\ g/mol}}$

With the problem clearly described, the numbers clearly entered, it is time to check your work by checking the units. This is a form of dimensional analysis. If your units don't come out right, then something went wrong.

To check this, cross out the units that cancel out in the numerator and denominator. In this case, grams cancels with grams and moles is left in the denominator of a denominator (This means it goes to the numerator. Check your algebra books for this if this confuses you.)

Mass = 120.0 g	Molar Mass = mass/moles
Molar Mass = 23.0 g/mol	Moles = mass/molar mass
n (moles) = ??	**Moles =** $\mathbf{\frac{120.0\ \cancel{g}}{23.0\ \cancel{g}/mol}}$

Finally, give your answer to the correct number of significant figures (in this case, 4 based on the measurement given in the original problem) and the correct unit.

Mass = 120.0 g	Molar Mass = mass/moles
Molar Mass = 23.0 g/mol	Moles = mass/molar mass
n (moles) = ??	Moles = $\frac{120.0\ \cancel{g}}{23.0\ \cancel{g}/mol}$ **Moles = 5.217391304347 = 5.217**
moles Na	

Often times, units should include substances. Think logically on these counts. If you say "5.220 moles," the question is 'moles of what?' Say moles of Sodium or "mol Na" to be clear.

<u>SHOW YOUR WORK FAQ</u>

Q: Do I have to show my work all the time?
A: When there is math or conversions involved, yes, it is appropriate to show your work.

Q: If I don't, can I lose points?
A: Frequently, and this also goes for work that is not coherent and clear. Don't make a grader search for the answer.

Q: What if that's how I solve a problem?
A: Unfortunately, SHOW YOUR WORK doesn't include the following:
· Cross multiplying. This is not work, it's unsolved algebra problems

· Long division or addition/subtraction/multiplication that is written out. Use a calculator for these.
Show me what the operation is neatly and then grab the calculator.
· A mess of numbers and lines that Pablo Picasso couldn't make sense of. Just writing it on the
page doesn't count. Again: Don't make the grader search for the answer.

Q: Is this always how I should show my work?
A: Different teachers may expect different things from students, but this is the clearest and most evident way of showing your thought process, so you should get used to it.

Q: Should every number have a unit?
A: Yes. Always.* A number without a unit is nothing.

*There are exceptions to this rule, but you will be directed to when this is the case.

SECTION 2 – DESIRED QUALITIES OF A SAINT MARY STUDENT

! Intelligence
This quality is not just about being "smart". It is being "smart" enough to identify what you do not know or understand and then actively seeking sources of help. This also includes knowing when you "get it", and when you need to stay after/ask for help.

! Self-Motivation
This quality describes your attitude. Enrollment in this class is required. Your desire to learn the material should be your chief motivation. You understand that the teacher will not cajole, plead, beg, etc. a student to do the assigned work. You should be ready and willing to learn each day.

! Integrity / Character
This quality is about doing the right thing in all situations. If you have integrity, you do not cheat on any assignment, be it a test, quiz, project or homework. You do your own work. If you have integrity it means you do not help others to cheat, be it providing homework for someone to copy or providing the questions / answers for a test or quiz in class or for another class.

! Work Ethic / Industriousness
This quality means that the work you turn in is of your highest quality. You show complete and organized work on all assignments (tests, quizzes, homework, projects) clearly identifying how you arrived at the solutions. Showing just answers does not show any work ethic at all and is unacceptable. Industriousness means that you use all available time to learn and improve. This could simply be starting your homework if there is time left in class. It could mean asking questions about a concept of which you are unsure. When given an extended problem / project / reading assignment industriousness means that you start on the assignment promptly and not wait until the night before the test or due date. This quality means you do not do work for another class or play games on your calculator during class time.

! Safety
Honors students treat the lab and lab materials with respect. While they may not yet know all the safety regulations, they do know that horsing around or misbehaving in the lab can potentially cause injury or worse to themselves and their peers. Chemistry students do not need to be told how to behave properly in a lab, or when to appropriately observe safe and correct lab techniques. Honors students ensure the lab is cleaner than when they found it. Labs should be read, at a minimum, the night before. You should highlight and write notes on your procedure. All prelab assignments should be done promptly and if there are questions you should discuss those with Mr. Lopez BEFORE the class period in which you are supposed to perform the lab.

! Inquisitiveness
This quality means that if you have a question you ask the question as soon as possible. A student does not just sit there and take notes, they think: Did I understand? Does it make sense? What if? Do not make the mistake of assuming that a concept you do not understand now in class will all make sense later on. Being inquisitive also means taking advantage of all opportunities to help yourself including:

Your teacher in class
Your teacher out of class
Your textbook!
Other students who may have a grasp of the concept

! Ingenuity
This quality is about applying knowledge, not just rote memorization. An honors student is able to devise solutions to problems they have never seen before. They are able to take what they have cumulatively learned in this class and all of their current and previous classes and apply it toward the solution of a new problem.

SECTION 3 – INTRODUCTION TO CHEMISTRY

Matter is anything that has mass and occupies space.

Chemistry is the study of the composition of matter and the changes that matter undergoes. Because living things are made of matter, chemistry affects all aspects of life and most natural events. Chemistry can explain how some creatures survive deep in the ocean where there is no light, or who some food tastes sweet and some taste bitter. It can even explain why there are different shampoos for dry or oily hair.

There are 5 divisions of chemistry, and they overlap somewhat.:

1. **Organic chemistry**: the study of substances containing carbon – often living, or once living, things
2. **Analytical chemistry** : the study of the composition of substances – how much of a chemical or substance is present
3. **Physical chemistry**: the study involving prediction of the behavior of chemicals – why does a substance do what it does under specific conditions or just in general
4. **Biochemistry**: the study of the chemistry of living organisms – how and why the chemistry of living organisms works and affects lives
5. **Inorganic chemistry**: the study of substances not containing carbon – nonliving things or things that never lived, and how they interact or react.

An integral part of chemistry is research. **Pure chemistry** research is the pursuit of chemical knowledge for its own sake. The chemist doesn't expect that there will be an immediate practical use for the knowledge. **Applied chemistry** research is directed toward a practical goal or application.

The **scientific method** is a logical, systematic approach to the solution of a scientific problem. Although the exact steps of the scientific method vary widely a rough requirement includes making observations, testing hypotheses, and developing theories.

In order to determine the appropriate approach to a problem, certain parameters must be defined:

The system
The relationship
The data

In the **system**, the actual boundaries of what are being observed is identified. A system can be anything from the contents of a test tube to an entire ecosystem. Some even consider the universe to be a system- albeit a large and unwieldy one, but one for making large generalizations. The system is the place in which you make your observations. If science were a story, the system would be the setting.

The **relationship** defines the actual cause and effect one is trying to find. The relationship is defined by a **hypothesis**, which is a testable (and disprovable) statement. Most hypotheses are "if-then" statements: "If I eat a lot of candy, then I will get sick." or "If I introduce a flame to diethyl ether, then I will get an explosion." Hypotheses must be testable, and the test for a hypothesis must leave an option that the hypothesis can be proven wrong. The test for a hypothesis will involve two kinds of information: variables and controls.

The **data** of the experiment is encapsulated by information that is controlled and

information that is varied. Controls are anything about the experiment that must be held steady.
A controlled environment is one where certain properties (temperature, air pressure, motion, composition of substances) are held constant throughout experimentation. This enables one to see the relationship between two variables much clearer and without the possibility of other factors affecting change.

Data in the form of either controls or variables is either quantitative or qualitative. Quantitative data is anything that is recorded as a number: temperature, volume, mass, etc. Qualitative data is anything that is recorded as a non-numerical observation: state of matter, color, texture, etc.

Observations: Something needs explanation – a problem, a question, an issue.
Hypothesis: A suggestion which solves the problem, answers the question, explains the issue is offered. It may be termed an "educated guess" because it is based on some logical explanation or suggestion of cause. Often stated as "If ____(independent variable)_______, then _____(dependent variable)______.
Experimentation: The testing of the hypothesis, usually repeated trials are necessary to Assure that the results of the tests can be accepted as genuine.
Independent Variable (Manipulated Variable): variable that is being controlled and experimentally changed
Dependent Variable (Responding Variable): variable which is observed during experiment
Theory: A theory, or explanation of the results of the experiment is offered. The theory is a possible answer, it cannot be proven to be true; it is possible to disprove a theory.
Scientific Law: A statement is offered which summarizes the results of experiments and observations. A scientific law does not try to explain the results, it merely states what the results are. The Law of Definite Proportions, the Law of Conservation of Mass. These laws are a result of observations. Theories may explain these laws, but the law is not an explanation. In essence, theories explain, laws describe.
Scientific Model: is a combination of theories into a coherent framework for use in explaining other scientific occurrences. For instance, this year, we will learn about how electrons, protons and neutrons interact and the theories of atomic structure. These theories are combined into a model of the atom, which will be used later on in the year when we will discuss how the atom interacts with other atoms to form chemical bonds. Models can be revised over time. And, in the case of the atom, we will learn how our own understanding of the atom has evolved with new information and new models have been developed because of this understanding.

SECTION 4 - THE METRIC SYSTEM/SI SYSTEM

In the next section, we introduce the standards for basic units of measurement. These standards were selected because they are reproducible and unchanging and because they allow us to make precise measurements. The values of fundamental units are arbitrary (Prior to the establishment of the National Bureau of Standards in 1901, at least 50 different distances had been used as "1 foot" in measuring land within New York City. Thus the size of a 100 ft by 200 ft lot in New York City depended on the generosity of the seller and did not necessarily represent the expected dimensions.)

In the United States, all units are set by the National Institute of Standards and Technology, NIST (Formerly the National Bureau of Standards, NBS). Measurements in the scientific world are expressed in the units of the METRIC SYSTEM or its modernized successor, the International System of Units (SI). The SI, adopted by the National Bureau of Standards in 1964, is based on the seven fundamental units listed in Table 1 below. All other units of measurement are derived from them.

Table 1 – The Seven Fundamental Units of Measurement (SI)

Physical Property	**Name of Unit**	**Symbol**
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	Kelvin	K
Amount of substance	mole	mol
Electric current	ampere	A
Luminous intensity	candela	cd

In chemistry we use the first five most often. The metric and SI systems are *DECIMAL SYSTEMS*, in which prefixes are used to indicate fraction and multiples of ten. The same prefixes are used with all units of measurement. The commonly used prefixes follow.

METRIC PREFIXES

	PREFIX	SYMBOL	MULTIPLE
BIG	Tera-	T	1,000,000,000,000.
	*giga -	G	1,000,000,000.
	*mega -	M	1,000,000.
	*kilo -	k	1,000.
	* hecto -	h	100.
	*deka -	da	10.
	meter	m	1
	gram	g	1
	liter	L	1
	*deci -	d	1/10
	*centi -	c	1/100
	*milli -	m	1/1000
	*micro	μ	1/1,000,000
	*nano -	n	1/1,000,000,000
SMALL	pico -	p	1/1,000,000,000,000

*These are the prefixes you should learn because you will encounter them most often.

TEMPERATURE

We sense temperature as a measure of the hotness or coldness of an object. Indeed, temperature determines the direction of heat flow. Heat always flows spontaneously from a substance of higher temperature to one at lower temperature. Thus we feel the influx of energy when we touch a hot object, and we know that the object is at a higher temperature that is our hand.

The temperature scales commonly employed in scientific studies are Celsius and Kelvin scales. The **Celsius scale** is widely used in chemistry and as the everyday scale of temperature in most countries. It is based on the assignment of 0 C to the freezing point and 100 C to its boiling point at sea level.

The **Kelvin scale**, however, is the SI temperature scale, and the SI unit of the temperature is the Kelvin (K). Historically, the Kelvin scale is based on the properties of gases. Zero on this scale is the lowest attainable temperature, -273.15 C, a temperature referred to as **ABSOLUTE ZERO**. Notice we do not use a degree sign for Kelvon. Both Celsius and Kelvin scales have equal-sized units- that is, a Kelvin is the same size as a degrees Celsius. Thus, the Kelvin and the Celsius scale are related as follows:

$$K = C + 273.15$$

The common temperature scale in the United States is the **Fahrenheit scale**, which is not generally used in scientific studies. On the Fahrenheit scale water freezes at 32 F and boils at 212 F. The Fahrenheit and Celsius scales are related as follows:

$$C = 5/9(F-32)$$

Or

$$F = 9/5(C)$$

Non SI Units

In our daily lives we usually do not utilize the SI Units. Conversions between non-SI and SI units are usually straightforward.

ENGLISH-METRIC EQUIVALENTS		
1 inch = 2.54 cm	1 pound = 453.6 g	1 quart = 0.9463 L

COMMON NON-SI UNITS			
Volume	liter	L	$1000\text{ L} = 1\text{ m}^3$ $1\text{ mL} = 1\text{ cm}^3 = 1\text{ cc}$
temperature	degrees Celsius	°C	K = °C + 273.15
heat energy	calorie	cal	1 cal = 4.184 J

For instance: Kilometer means "1000 meters," in the SI system 'kilo' means 1000 or 10^3. If I have 54 meters, I have

$$54\ \cancel{m} \times \frac{1\text{ km}}{1000\ \cancel{m}} = 0.054\text{ km}$$

If I have 20 kilometers, I have

$$20\ \cancel{km} \times \frac{1000\text{ m}}{1\ \cancel{km}} = 20\,000\text{ m}$$

This method of conversions is known as Dimensional Analysis and is covered in the next section of this packet.

SECTION 5 – DIMENSIONAL ANALYSIS

Dimensional analysis is a way of examining the measurements of a problem to determine the answer to the problem.

For instance: How many quarters are in 12 dollars?

You'll probably be able to figure this one out in your head. Or at least your instinct is to grab a calculator, but before you do, we'll need to solve this problem with dimensional analysis. This problem may be easy to solve in your head, but others you'll do in Chemistry will require this method of solving to give you the correct answer (or to be counted in "Show Your Work" – see attached paper titled SHOW YOUR WORK to understand what this phrase means.) First, to do a conversion, an equality must be known that compares quarters and dollars. In this case, we know that 1 dollar equals 4 quarters. Write this algebraically as:

$$1 \text{ dollar} = 4 \text{ quarters}$$

This is the equality that will be turned into a conversion factor. Either divide both sides by "1 dollar" to get

$$\text{(A)} \quad 1 = \frac{4 \text{ quarters}}{1 \text{ dollar}}$$

Or divide both sides by "4 quarters" to get

$$\text{(B)} \quad 1 = \frac{1 \text{ dollar}}{4 \text{ quarters}}$$

Notice that both these conversion factors equal 1. Conversion factors always equal 1. They are not for changing numbers, conversion factors are for changing units. If your conversion factor does not equal 1, it means that the equation you derived it from is not equal.

Using this conversion factor, we have to decide which to use (A) or (B). This is determined by the unit path we're taking. To get to 'quarters,' we have to get rid of 'dollars' and to do that we multiply our given value by the conversion factor that helps us get rid of 'dollars' and get to 'quarters'

Given	x	conversion factor	=	answer
(A) 12 ~~dollars~~	x	$\frac{4 \text{ quarters}}{1\ \cancel{\text{dollar}}}$	=	48 quarters
(B) 12 dollars	x	$\frac{1 \text{ dollar}}{4 \text{ quarters}}$	=	$3 \frac{\text{dollars}^2}{\text{quarter}}$

In (A), the dollars and dollars units are able to cancel out algebraically, leaving only the unit quarters, which is what is sought. So that is the correct conversion factor. In (B), the wrong conversion factor is used. Notice how dollars and dollars can't cancel out. The unit given as a result is nonsensical. If the unit is nonsensical, then the answer cannot be 3, it must be 48. This makes sense logically as well. Don't believe me? Then let me trade you these 12 dollars for 3 quarters.

Conversion factors can be chained, which means that as long as the conversions are sensible and they are lined up correctly, multiple conversions can be lined up. This saves time when attempting to make many conversions. As long as units are maintained, you can trust this conversion method.

How many hours in 2 weeks?

Two things are known:

There are 7 days in 1 week
There are 24 hours in 1 day

There is no conversion factor easily known that converts hours to weeks (clearly there is, but you probably don't know it off the top of your head.) Instead, we will chain our conversions. First, start with what is given:

2 weeks

Then line up the conversion that matches 2 weeks and will cancel out the unit 'week.'

$$2 \text{ weeks} \times \frac{7 \text{ days}}{1 \text{ week}}$$

Instead of solving, continue on, by lining up the next conversion factor that cancels out days and gives you hours:

$$2 \text{ weeks} \times \frac{7 \text{ days}}{1 \text{ week}} \times \frac{24 \text{ hours}}{1 \text{ day}} =$$

The units cross out and the only unit left is the one you wish to find:

$$2 \cancel{\text{weeks}} \times \frac{7 \cancel{\text{days}}}{1 \cancel{\text{week}}} \times \frac{24 \text{ hours}}{1 \cancel{\text{day}}} =$$

Try the conversions in the question section. These conversions are simple conversions. Remember that every number MUST have a unit attached to it, in every step of the work you show. No points will be given if no work is shown.

You should line up your work as shown above when asked to show your work for conversions.

SECTION 6 - SCIENTIFIC NOTATION

In chemistry, we measure and calculate many things, so we must be sure we understand how to use numbers. In this section we will discuss the notation of very large and very small numbers.

We use **scientific notation** when we deal with very large and very small numbers. For example, 197 grams of gold contains approximately

602,000,000,000,000,000,000,000 gold atoms

The mass of one gold atom is approximately

0.000 000 000 000 000 000 000 327 grams

In using such large and small numbers, it is inconvenient to write down all the zeroes. In scientific (also known as exponential) notation, we place one nonzero digit to the left of the decimal.

$$4{,}300{,}000 = 4.3 \times 10^{6}$$

6 places to the left, therefore exponent of 10 is 6
If you move the decimal to the *LEFT*, the exponent is *POSITIVE!*

$$0.000348 = 3.48 \times 10^{-4}$$

4 places to the right, therefore exponent of 10 is -4
If you move the decimal to the *RIGHT*, the exponent is *NEGATIVE!*

The reverse process converts numbers from exponential to decimal form. In scientific notation the digit term indicates the number of significant figures in the number. The exponential term merely locates the decimal point and **DOES NOT** represent significant figures.

Realize that in general, if the number you're using is less than 1, the scientific notation exponent should be negative; if the number is greater than 10, the scientific notation exponent should be positive.

To convert back from scientific notation, simply move the decimal back the appropriate number of places.

5.3x104 becomes 53000. when the decimal is moved 4 places to the right.
4.9x10-4 becomes 0.00049 when the decimal is moved 4 places to the left.

The University of Maryland has a website for practicing Scientific Notation conversions at: http://janus.astro.umd.edu/cgi-bin/astro/scinote.pl

This game will help you learn to convert from scientific notation into standard notation and back again.

Putting scientific notation in calculators: Students often make mistakes when they try to enter numbers into their calculators in scientific notation. Below is the simplest and most foolproof way of doing this.

If you want to enter 4.36×10^{-2} into your calculator.

- Press 4.36
- Press EE or EXP, which stands for "times ten to the"
- Press -2 (the magnitude of the exponent)

The calculator display might show the value as 4.36^{-02} or as 0.0436. Either way is correct. Different calculators show different numbers of digits.

CAUTION: Be sure you remember that the EE or EXP button *includes* the "times 10" operation. An error that beginners make is to enter "x 10" explicitly when trying to enter a number in scientific notation. This will cause an error in **all of your calculations.**

1. Addition and Subtraction

In addition and subtraction all numbers are converted to the same power of 10, and the digit terms are added or subtracted.

$4.21 \times 10^{-3} + 1.4 \times 10^{-4} \rightarrow 4.21 \times 10^{-3} + 0.14 \times 10^{-3} = \mathbf{4.35 \times 10^{-3}}$

$8.97 \times 19^{4} - 2.31 \times 10^{3} \rightarrow 8.97 \times 10^{4} - .231 \times 10^{4} = \mathbf{8.74 \times 10^{4}}$

2. Multiplication

The digit terms are multiplied in the usual way, the exponents are added algebraically, and the product is written with one nonzero digit to the left of the decimal.

$4.7 \times 10^{7} \times 1.6 \times 10^{2} = 4.7 \times 1.6 \times 10^{7+2} = \mathbf{7.5 \times 10^{9}}$

$8.3 \times 10^{4} \times 9.3 \times 10^{-9} = 8.3 \times 9.3 \times 10^{4+(-9)} = \mathbf{7.7 \times 10^{-4}}$

3. Division

The digit term of the numerator is divided by the digit term of the denominator, the exponents are subtracted algebraically, and the quotient is written with one nonzero digit to the left of the decimal.

$$\frac{8.7 \times 10^{7}}{2.0 \times 10^{3}} = \frac{8.4}{2.0} \times 10^{7-3} = \mathbf{4.2 \times 10^{4}}$$

$$\frac{3.81 \times 10^{9}}{8.412 \times 10^{-3}} = \frac{3.81}{8.412} \times 10^{9-(-3)} = 0.45292 \times 10^{12} = \mathbf{4.53 \times 10^{11}}$$

4. Powers of Exponentials

The digit term is raised to the indicated power, and the exponent is multiplied by the number that indicates the power.

$(1.2 \times 10^3)^2 = 1.2^2 \times 10^{3 \times 2} = 1.44 \times 10^6 = \mathbf{1.4 \times 10^6}$

$(3.0 \times 10^{-3})^4 = 3.0^4 \times 10^{-3 \times 4} = 81 \times 10^{-12} = \mathbf{8.1 \times 10^{-11}}$

5. Roots of Exponentials

The exponent must be divisible by the desired root if a calculator is not used. The root of the digit is extracted in the usual way, and the exponent is divided by the desired root.

$(2.5 \times 10^5)^{1/2} = (25 \times 10^4)^{1/2} = 25^{1/2} \times (10^4)^{1/2} = \mathbf{5.0 \times 10^2}$

$(2.7 \times 10^{-8})^{1/3} = (27 \times 10^{-9})^{1/3} = 27^{1/3} \times (10^{-9})^{1/3} = \mathbf{3.0 \times 10^{-3}}$

SECTION 7 – ALGEBRAIC TRANSFORMS

In order to properly show work for calculations in Honors Chemistry (For a detailed description see SHOW YOUR WORK page in this packet) you will need to be able to transform an algebraic equation to solve for any given variable you are trying to find.

$x = yz$

In order to solve for 'z', both sides must be divided by 'y' to get

$z = y/x$

The same happens for

$xy = zab$

This equation solved for "a" requires both sides to be divided by "zb"

$a = xy/zb$

If squaring a term, recall that the opposite of a square is a square root. So for the equation

$x = y^2$

to solve for 'y', square root both sides:

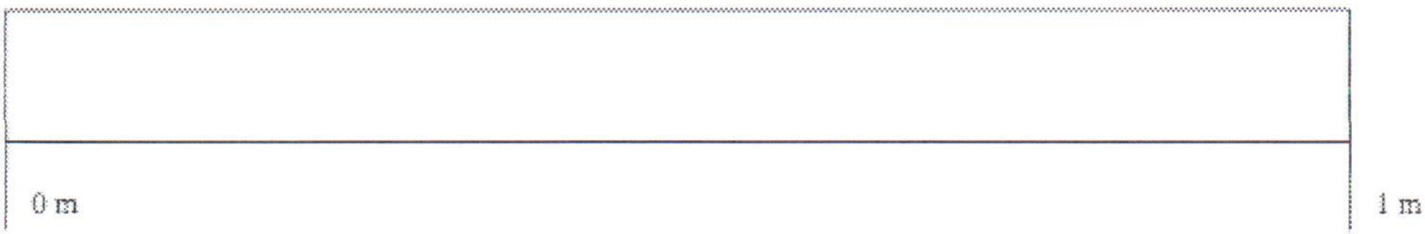
$y = \sqrt{x}$

SECTION 8 - SIGNIFICANT FIGURES/READING EQUIPMENT PROPERLY

SIGNIFICANT FIGURES MATTER WHEN YOU TAKE A MEASUREMENT. IF YOU DO NOT HAVE A MEASUREMENT, THEN THERE ARE NOT SIGNIFICANT FIGURES!

Significance in measurements is the degree to which the measurement was made. A more precise measurement can be made with a ruler than can be made with a meter stick. This translates to more numbers past the decimal.

Measuring a room with a meter stick may yield that the room is 5 m by 5 m, although you could measure the half-meter and you could say that the room is 5.3 m by 5.4 m. The last digit in this case is known as the estimated digit. The estimated digit is what is measured between the actual measurements on the measuring device. (Please note, rulers are not to scale.)



One can estimate between 0 and 1 meters on this stick so the last digit (or estimated digit) is the tenths of a meter place.

If one uses a meter stick that has divisions for decimeters, then one can measure more precisely to the tenths place, and can estimate between the division of the decimeter stick. This means that the estimated digit can go to the hundredths place. Instead of measuring a room to be 5.3 m x 5.4 m, one could measure the room to be 5.34 m x 5.39 m. The precision increases with an increase in digits.

In Science, it is important that precision is appropriately assigned. In order to do this, one must be able to count significant figures in a number.

5.5 is a number that has 2 significant figures, the ones place and the tenths place.

5.53 is a number that has 3 significant figures. It is more precise than a number that has only 2 significant figures.

When looking at a measured number, the number of figures given is typically the number of significant figures measured.

When calculating numbers using a calculator, often times false significance or precision can be achieved. This is when two numbers, multiplied or divided together yield a number that appears to be more precise (has more digits) than either of the measured values can actually determine.

A room is 2.5 m by 2.5 meters. What is the area of the room?

A calculator will tell you that the area is 6.25 m2, but this gives a degree of false precision. Your measurements were only precise to the 10ths place, but the answer yielded is precise to the hundredths place. Even though this is the exact answer a calculator will give, the number must be adjusted to suit the imprecision of the measurement:

Area = 6.3 m^2

This number is rounded to the correct number of significant figures allowed by the measured numbers.

When calculating numbers in science, you are bound in your answer by the measurement that is least significance. If you take two measurements of the room, one to be 2.5304 m and the other to be 2.52 m, the answer will have to have 3 significant figures.

Even though the answer will be 6.376608 m^2 when plugged into a calculator, one of your measurements only has 3 significant figures. This means your answer can only have 3 significant figures.

Area = 6.38 m2

It is important to be honest when reporting a measurement, so that it does not appear to be more accurate 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.

1. Determining the Number of Significant Figures in a Number

The number of significant figures in a measurement, such as 2.531, is equal to the number of digits that are known with some degree of confidence (2, 5, and 3) plus the last digit (1), which is an estimate or approximation. As we improve the sensitivity of the equipment used to make a measurement, the number of significant figures increases.

Postage Scale	3 g	?1 g	1 significant figure
Two-Pan Balance	2.53 g	?0.01 g	3 significant figures
Analytical Balance	2.531 g	?0.001 g	4 significant figures

Thus, we use significant figures to show just how sure we are about a measurement. Everything we measure has some uncertainty no matter if we are using a ruler, a graduated cylinder, or even looking at the speedometer of our car. There is no perfect measurement tool so we need to be responsible in reporting them.

2. Rules for counting significant figures in numbers.

- Any non zero digits are **always** significant.
- Zeros *within* a number are **always** significant. Both 4308 and 40.05 contain four significant figures.
- Zeros that do nothing but hold the place for a decimal point are **not** significant. Thus, 470,000 and 0.0048 both have two significant figures.
- Zeros at the end of a number that contains a decimal point are significant (think of it as we wouldn't write them unless they were important). For example, 4.00 has three significant figures and 4500. Has four significant figures.

Examples:

435 = 3 s.f.	*all digits significant*
0.02304 = 4 s.f.	*lead in zeros don't matter they are just holding spots*
980,000 = 2 s.f.	*trailing zeros don't matter because there is no decimal point*
4.509 x 10^2 = 4 s.f.	*all numbers in a scientific notation are significant.*

3. Addition and Subtraction with Significant Figures

When adding numbers, the accuracy of the final answer can be no greater than the least accurate measurement. SIMPLY STATED: When measurements are added or subtracted, the answer has the least number of decimal places.

Examples:

150.0 + 0.507 = 150.5 *because 150.0 only has one decimal so the answer has 1*

235.78 – 34.907 = 200.87 *because 235.78 has two decimals so the answer has 2*

Be sure to round your answers correctly if necessary

4. Multiplication and Division with Significant Figures

The same principle governs the use of significant figures in multiplication, however, we count the total number of significant figures in each measurement, not the number of decimal places. SIMPLY STATED: When measurements are multiplied and divided, the answer has the least number of TOTAL significant figures.

Examples:

34.5 x 6.0 = 207 . . .(rounded) = 210 *6.0 has two sig figs so answer has 2 sig figs*

808.6/2.3 = 351.5652174 (rounded) = 350 *2.3 has two sig figs*

5. More than Step Math

When you have more than one mathematical operation you need to do sig figs for each step of the problems.

Examples: (45.61 + 2.3) x 12

Step 1: 45.6 + 2.3 = 47.91 (rounded) = 47.9

Step 2: 47.9 x 12 = 574.9 (rounded) = 570 FINAL ANSWER HAS 2 Sig Figs

In order to practice, you can use an applet produced by California State University to practice determining significant digits.

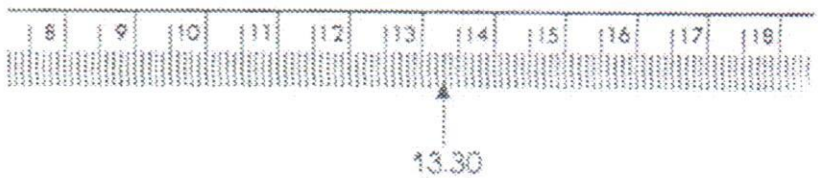
http://chemistry2.csudh.edu/lecture_help/sigfigures.html

If you need additional explanation on these concepts, search online. Rensselaer Polytechnic Institute has this site to explain significant figures. It also has exercises you can use to practice performing calculations with significant figures.

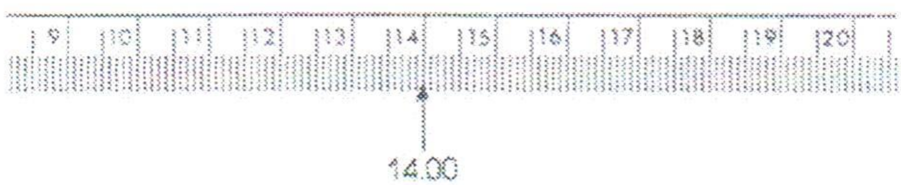
http://www.rpi.edu/dept/phys/Dept2/APPhys1/sigfigs/sigfig/node10.html

USING EQUIPMENT

On an ordinary ruler, the smallest divisions are millimeters, 0.01 m. Rarely does the object being measured end very neatly at one of the lines of the instrument. However since this does happen occasionally and forces the use of measurement (and not place-holder) zeros, let's deal with that first.



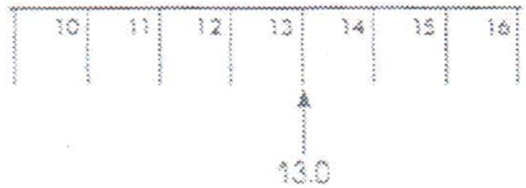
To indicate that the object being measured ends *EXACTLY* at the third line after the 13, we *MUST* write 13.3**0** cm. This indicates that, to the best of our ability, the measurement does not extend into the hundredth of the centimeter. This number could also be expressed as 0.1330 m or 133.0 mm. Notice that the number of significant figures does not change even when we convert it into a new measurement.



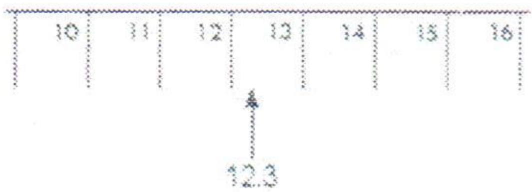
Now, according to our best estimation, the object ends *EXACTLY* at the 14 cm line. Now we must have two zeroes at the end. The number must be written at 14.00 cm. This indicates that the measurement does not extend into the tenth of a centimeter space and that the instrument is precise to an estimation in the hundredth of centimeters. Of course, the number could also be expressed as 0.1400 m or 140.0 mm. The number must end with two zeros and notice that the number of significant figures does not change when we convert.

Most of the time, our measurement falls between the lines and, making sure that the instrument is directly in front of our eyes, we must make agonizing estimates about where the measurement does fall. We could estimate the above reading to be 12.85 cm. This means that we "guess" that the object falls midway between 12.8 cm and 12.9 cm. This could also be written as 0.1285 m or 128.5 mm.

All meter sticks are not created equal. Suppose we come upon one with no millimeter markings:



This reading would be 13.0 cm or 0.130 m or 130. mm . One zero only can and must be written. It indicates that there are no tenths of centimeters on the meter stick.



We could estimate the reading about to be 12.3 cm. We must write the 3 but may not write any more figures.

Remember that the last figure written in any measurement is always an estimate, so that the last figure is always doubtful. There may be one and **only one** estimated or doubtful figure in any measurement.

Rules for Reading Any Instrument

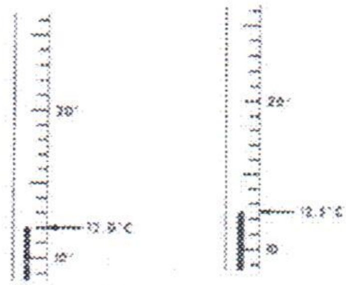
The precision of the instrument determines the number of significant figures in the measurement. Ordinarily, the instrument should be read to its limit. The last figure in any measurement is doubtful.

On these metric rules, the estimate is between the millimeter lines.

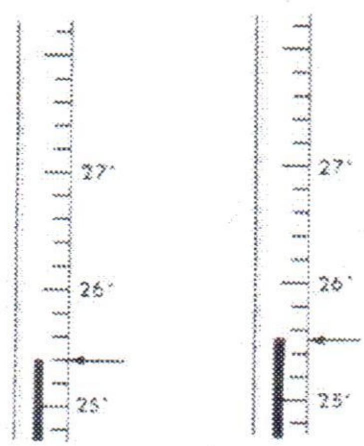
The numbered lines represent the centimeter lines.

THERMOMETERS

Most glass laboratory thermometers used by students who are just beginning to learn about significant figures are marked with single degrees as the most basic scale. Estimations, then, are usually in tenths of degrees. Very frequently, the space between the lines is so small that the only practical estimation is exactly on the degree line or halfway between degree lines.



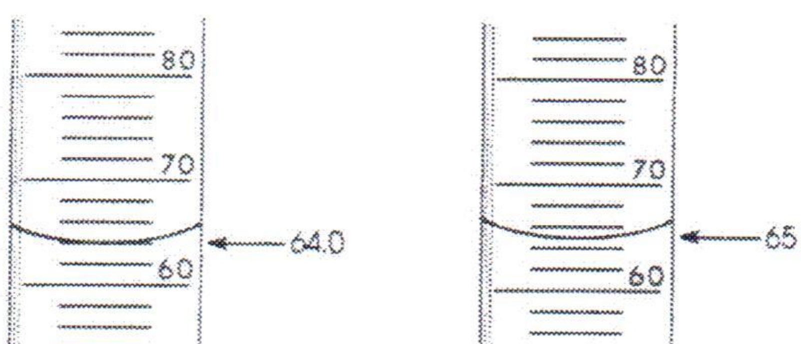
Even though these thermometers are the most common in school labs, the students are not excused from checking the scale before making a reading. Notice the following.



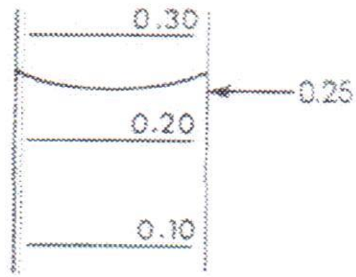
As when reading *any* scale, you must be sure that the thermometer is at eye level and directly in front of you when reading. The last and *only the last* figure of the reading is an estimate.

GRADUATED CYLINDERS

You should be able to measure the volume of liquids in a graduated cylinder. How precisely you can measure volume depends on the size and type of graduated cylinder you use. Generally, you should be able to estimate between etched or printed lines.



When the gradations are 0.2 mL apart, as in the above graduates, we write a figure in the tenth milliliter place if we estimate the meniscus to rest exactly on the line. If the meniscus falls between lines, however, we may only write to the nearest whole number.



Graduates are very often have gigantic liter sizes with very big spaces between the markings. Of course, the bigger the spaces, the less dependable the estimate.

REMEMBER:

All readings must be made at eye level.

We read the level of the liquid at the bottom of the meniscus.

The calibrations on graduated cylinders vary widely. You must determine the precision on each shown on the next page and read it accordingly. All numbers represent milliliters.

Chemistry

ANSWER BOOKLET – Summer Packet

Answers and work must be recorded in or attached to this packet. NO EXCEPTIONS

SECTION 3 – INTRODUCTION TO CHEMISTRY

1. Which division or divisions of chemistry might be used to examine:
   a) The mechanism by which blood clots
   b) The amount of a toxic substance found in a water supply
   c) The reason a metal melts at a specific temperature
   d) The formation of chemical compounds made up of metals

Casey noticed that the plants near the river downstream from the local copper mine were withered and dying. She wanted to find out if there was something from the mine that was causing this. She set out to measure levels of water from the river for concentrations of copper, lead, arsenic and sulfates, recorded these levels and compared them to levels of water from the river upstream from the mine. Casey then sought to check the different metal concentrations and their effect on plant growth by measuring the height of plants grown watered with different solutions of copper, lead, arsenic and sulfates. She also included a group of plants that were grown with regular water, untreated with chemicals. She recorded this information over the course of 3 months and also included observations such as plant color, and leaf texture. When she was done, she compiled her data and generated plots and graphs to see if she could determine any relationships between the different toxins and plant growth. Casey noticed that there seemed to be a steep decrease in the height of plants grown with water and lead, along with water and copper. Water with sulfates showed a decline in plant height, but not as steeply. Plants watered with arsenic did not grow at all.
Casey performed an experiment here that was based upon a scientific method. Answer the following questions based on what you know of the scientific method and from the passage above.

2. Although a hypothesis was not overtly stated by Casey, which of the following would be the most appropriate hypothesis? (Circle the appropriate letter)

a. If plants are grown downriver from a copper mine, then they will die.
b. If one checks river water, then one will find traces of lead, copper and other metals.
c. If plants are watered with water containing trace concentrations of certain metals, then there will be a decrease in their plant height.
d. If copper mines are near rivers, then copper will be found in the water.

3. Although not all of the controls were mentioned from Casey's experiment, think about performing this experiment yourself. What are 3 controlled variables that you think were involved in this experiment?

a. ____________________________ b. ____________________________ c. ____________________________

4. Identify the independent variable in this experiment and the dependent variable.

Independent –
Dependent –

5. Identify an example of quantitative data in this experiment and an example of qualitative data.

Quantitative –

Qualitative –

6. Write an example of a conclusion to this experiment that addresses both the relationship that was sought (the hypothesis) and the data was collected. This conclusion should not be more than 2 sentences, but should be written in complete sentences.

____

____

____

SECTION 4 – METRIC SYSTEM/SI SYSTEM – attach all work clearly and neatly for this section

7. Perform the following conversions:
   a. 454 mg to g
   b. 5.0×10^{-9}m to pm
   c. 3.5×10^{-2} mm to μm
   d. 36.3 km to m
   e. 447 kg to g
   f. 55.9 dL to L
   g. 6251 L to cm^3

   <u>Conversion for (b)</u> 1 pm = 1.0×10^{-12} m

8. Make the following temperature conversions:
   a. 233 C to K
   b. The melting point of potassium iodide is 681 C. What is this temperature in Kelvins?

9. Perform the following conversions:
   a. 8.60 miles to m
   b. 3.00 days to s
   c. 16.2 ft to m

SECTION 5 – DIMENSIONAL ANALYSIS – attach all work clearly and neatly for this section

Show, using dimensional analysis, how to convert from the given unit to the sought unit. [Answers] are provided. This assignment is checked to see if work can be done properly. Remember, every number MUST have a unit.

10. How many feet are in 10 meters? (2.54 cm = 1 inch) [32.8]
11. How many dozens of doughnuts are 144 doughnuts? [12]
12. How many seconds are in 1 year? (1 year = 365 days) [31 536 000]
13. How many quarters are in $43.75? [175]
14. How many nickels are in that same amount? [875]
15. How many dollars are in 98 quarters? [24.50]
16. How many months are in 17 500 000 seconds? (1 month = 30 days) [6.75]

SECTION 6 – SCIENTIFIC NOTATION

17. Perform the following calculations.
   a. $4.0 \times 10^{15} - 3.0 \times 10^{14}$
   b. $15 \times 10^{-5} + 6 \times 10^{-6}$
   c. $9 \times 10^{7} - 20 \times 10^{6} + 10 \times 10^{7}$
   d. $4 \times 10^{13} \times 3 \times 10^{17}$
   e. $8 \times 10^{54} / 2 \times 10^{32}$
   f. $6 \times 10^{79} \times 6 \times 10^{1}$
   g. $42 \times 10^{100} / 7 \times 10^{65}$
   h. square root of 144×10^{16}

18. Complete the following chart.

Standard Notation	Scientific Notation
104 000 m	1.04×10^{5} m
0.000 543 g	5.43×10^{-4} g
0.004 5 mol	
60 200 000 s	
	3.403×10^{7} L
	1.22×10^{-3} J
16 700 kPa	
150 W	
1 000.30 m^3	
602 000 000 000 000 000 000 000 particles	

SECTION 8 – SIGNIFICANT FIGURES/READING EQUIPMENT PROPERLY

19. How many significant figures are in each of the following numbers (assume that each number is a measured quantity):

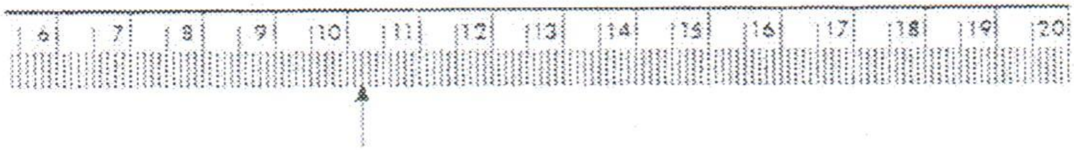
a. 4.003
b. 6.023×10^{23}
c. 5000
d. 1282 kg
e. 0.00296 s
f. 8.070 mm
g. 0.0105 L
h. 9.7750×10^{-4} cm
i. 0.0234 m^2
j. 7,194,300 cm
k. 435.983 K
l. 204.080 g

20. Carry out the following operations, and express the answers with the appropriate number of significant figures.

a. 1.24056 + 75.80
b. 23.67 – 75
c. 890.00 x 112.3
d. 78,132/2.50
e. 37.24 mL + 10.3 mL
f. 21.2342 g – 27.87 g
g. 1.23 cm x 12.34 cm
h. 906.34 – (8903.2/5.7)
i. 356.2 x 10^{4}- 2.4 x 10^{3}) x 3.97
j. 482 x [2537 – (3.76 x 90)]

21.

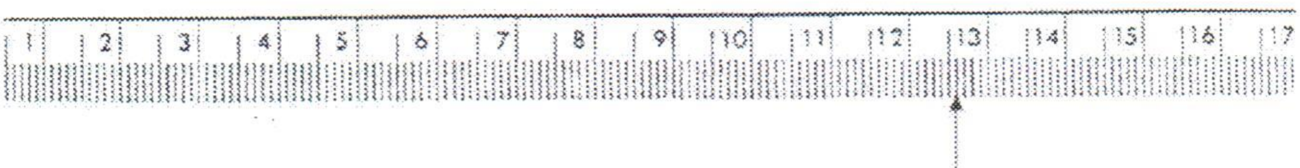
22.


6 7 8 9 10 11 12 13 14 15 16 17 18 19 20

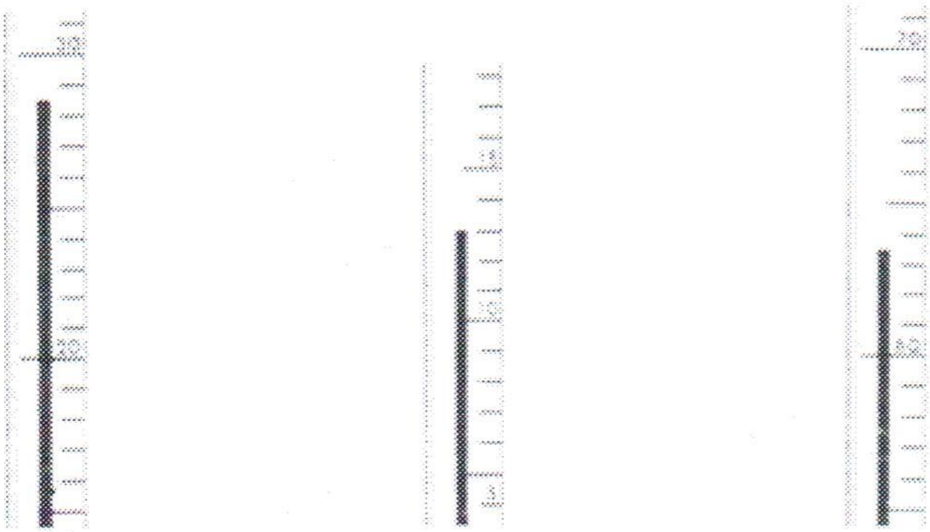
23.


1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17

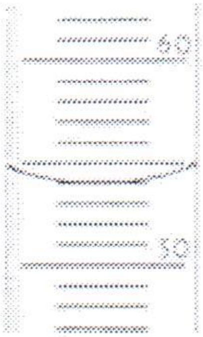
24.


1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17

25.



26.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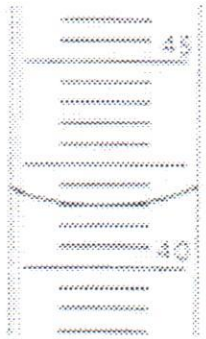
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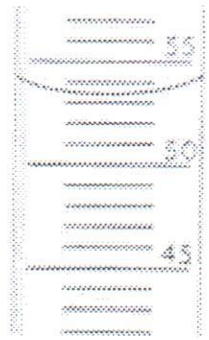
55
50
45

SECTION 7 – ALGEBRAIC TRANSFORMATIONS

Density Equation	Light Equation	Energy of Light Equation	Combined Gas Law	Ideal Gas law	Molar Mass Determination	Graham's Law of Effusion
$\rho = \frac{m}{V}$	$c = \lambda v$	$E = hv$	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$	$PV = nRT$	$M = \frac{m}{n}$	$\frac{R_A}{R_B} = \frac{\sqrt{M_B}}{\sqrt{M_A}}$
$\rho=$	$c=$	$E=$	$P_1=$	$P=$	$M=$	$R_A=$
$m=$	$\lambda=$	$h=$	$P_2=$	$V=$	$n=$	$R_B=$
$V=$	$v=$	$v=$	$T_1=$	$n=$	$m=$	$M_A=$
			$T_2=$	$R=$		$M_B=$
			$V_1=$	$T=$		
			$V_2=$			