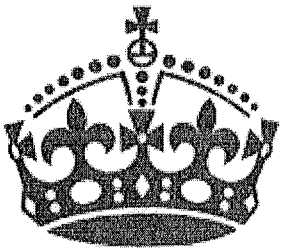
Name_____ Mr. GARDNER

Unit #1: Chemistry Introduction/Matter and Energy

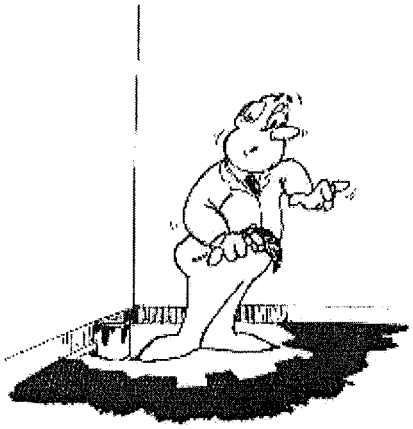


KEEP CALM AND STUDY CHEMISTRY

Index

STUDY TIPS: THINGS TO ASK YOURSELF

1) Do I have all the notes?
2) Have I spent at least 5-10 minutes each day reviewing what was discussed in class that day as well as going over previous material?
    a. Re-read your notes many times
    b. Do I understand what the notes mean?
        i. If confused, write down a question and bring it to class or after school review
3) Have I completed all the homework?
    a. Am I doing it during the commercials of my favorite show?
    b. Am I copying it from my friend without really understanding it?
    c. Do I correct my mistakes when we go over answers/solutions in class?
4) Do I ask questions in class or am I too embarrassed, so that I remain unsure of the material?
5) Have I read the text book chapter to support material covered in the notes and lectures?
6) Have I gone online for additional review help?
7) Have I put in a satisfactory effort that is good enough for me?
    a. Are you proud of your efforts?
    b. How has past work paid off or resulted?



Don't paint yourself into a corner

*****YOU GET OUT OF IT WHAT YOU PUT INTO IT*****

Key Objectives: Chem Intro **GARDNER**

- Define the terms chemistry, matter, pure substance, and mixture.
- Distinguish among elements, compounds, and mixtures.
- Distinguish between physical and chemical properties.
- Distinguish between physical and chemical changes.
- List the common metric units used in chemistry.
- Name the most commonly used metric prefixes and their numerical equivalents.
- Define the terms volume and density as they apply to chemistry.
- Describe how measurement readings are taken and estimated.
- Define and apply the terms accuracy, precision, and significant figures.
- Apply the rules for adding (subtracting) and multiplying (dividing) measurements.
- Define the term percent error, and calculate the percent error of a measurement.
- Understand the factor-label method (FLM), and use it to solve problems.

Key Objectives Formulas and Equations

- Define the term binary compound.
- Define the term oxidation number and assign oxidation numbers to elements, simple ions, and the elements contained in compounds and polyatomic ions.
- Write the formulas for binary compounds and for compounds containing polyatomic ions.
- Name binary compounds and compounds containing polyatomic ions using the Stock system.

Matter

condensation an exothermic process in which a vapor or a gas changes into the liquid phase; the potential energy of the substance decreases during this constant temperature process; the reverse of the vaporization process

deposition the process in which a gas changes directly into a solid; the reverse of sublimation

freezing the constant temperature process in which particles in the liquid phase lose energy and change into the solid phase; also known as solidification; the reverse of the melting process

fusion the constant temperature process in which particles in the solid phase gain enough energy to break away into the liquid phase; also known as melting; the reverse of the freezing process; (in nuclear chemistry) the combining of light nuclei into a heavier nucleus

gaseous phase a phase of matter without definite shape or volume

heat energy transferred from one substance to another; measured in units of calories or joules
heat of fusion the amount of heat needed to convert a unit mass of a substance from a solid to a liquid at its melting point

heat of vaporization the amount of heat needed to convert a unit mass of a substance from a liquid to a vapor at its boiling point

kinetic molecular theory a theory used to explain the behavior of gases in terms of the motion of their particles

liquid phase a phase of matter having definite volume but no definite shape (takes the shape of its container)

solid phase a phase of matter having a definite shape and volume; particles in this phase have a definite crystalline arrangement

sublimation the process in which a solid changes directly into a gas; the reverse of deposition
temperature the measure of the average kinetic energy of a substance's particles

vaporization the constant temperature process in which particles in the liquid phase gain enough energy to break away into the gaseous phase; also known as boiling; the reverse of the condensation process

space for additional terms:

Formulas and Equations

analysis a chemical reaction in which a compound is broken down (decomposed) into simpler substances.

chemical change a reaction in which the composition of a substance is changed.

coefficient the number placed before a formula indicating the number of units of that substance

decomposition a chemical reaction in which a compound is broken down into simpler substances

diatomic molecule a molecule containing two identical atoms

double replacement a chemical reaction in which ions exchange places

empirical formula the simplest integer ratio in which atoms combine to form a compound

endothermic a chemical reaction that absorbs heat, producing products with more potential energy than the reactants

exothermic a chemical reaction that releases heat, producing products with less potential energy than the reactants

formula symbols and subscripts used to represent the composition of a substance

molecular formula the actual ratio of the atoms in a molecule

molecule the smallest unit of a covalently bonded substance that has the properties of that substance

physical change a change that does not alter the chemical properties of a substance

polyatomic ion a covalently bonded group of atoms that have a net electric charge

product a substance formed in a chemical reaction, shown to the right of the arrow in an equation

qualitative information that cannot be counted or measured

quantitative information that can be either counted or measured

reactant a starting substance in a reaction, shown to the left of the arrow in an equation

single replacement a reaction in which an element replaces a less reactive element in a compound

subscript the number written after a chemical symbol in a formula indicating the number of atoms present

symbol a one-, two- or three-letter designation of an element

synthesis a reaction in which two or more sub- stances combine to form one product

space for additional terms:

BAN DIHYDROGEN MONOXIDE!

Dihydrogen monoxide is colorless, odorless, tasteless, and kills uncounted thousands of people every year. Most of these deaths are caused by accidental inhalation of DHMO, but the dangers of dihydrogen monoxide do not end there.

Prolonged exposure to its solid form causes severe tissue damage. Symptoms of DHMO ingestion can include excessive sweating and urination, and possibly a bloated feeling, nausea, vomiting and body electrolyte imbalance. For those who have become dependent, DHMO withdrawal means certain death.

Dihydrogen monoxide:

- is also known as hydroxl acid, and is the major component of acid rain.
- contributes to the "greenhouse effect."
- may cause severe burns.
- contributes to the erosion of our natural landscape.
- accelerates corrosion and rusting of many metals.
- may cause electrical failures and decreased effectiveness of automobile brakes.
- has been found in excised tumors of terminal cancer patients.

Contamination is reaching epidemic proportions!

Quantities of dihydrogen monoxide have been found in almost every stream, lake, and reservoir in America today. But the pollution is global, and the contaminant has even been found in Antarctic ice. DHMO has caused millions of dollars of property damage in the midwest, and recently California.

Despite the danger, dihydrogen monoxide is often used:

- as an industrial solvent and coolant.
- in nuclear power plants.
- in the production of styrofoam.
- as a fire retardant.
- in many forms of cruel animal research.
- in the distribution of pesticides.
- as an additive in certain "junk-foods" and other food products.

Even after washing, produce remains contaminated by this chemical.

Companies dump waste DHMO into rivers and the ocean, and nothing can be done to stop them because this practice is still legal. The impact on wildlife is extreme, and we cannot afford to ignore it any longer!

The American government has refused to ban the production, distribution, or use of this damaging chemical due to its "importance to the economic health of this nation." In fact, the navy and other military organizations are conducting experiments with DHMO, and designing multi-billion dollar devices to control and utilize it during warfare situations. Hundreds of military research facilities receive tons of it through a highly sophisticated underground distribution network. Many store large quantities for later use.

Laboratory Equipment

In the chemistry laboratory, you will be using equipment, handling materials, and performing certain unfamiliar tasks. The purpose of this section is to introduce you to the equipment and to describe some of the skills you will be required to use in this laboratory course. A few of the "dos" and "don'ts" that are necessary for safe and effective laboratory work are also included.

RECOGNIZING LAB EQUIPMENT

The equipment you will be using most frequently in the laboratory is illustrated in Figure 1. Study this figure carefully and familiarize yourself with each item.

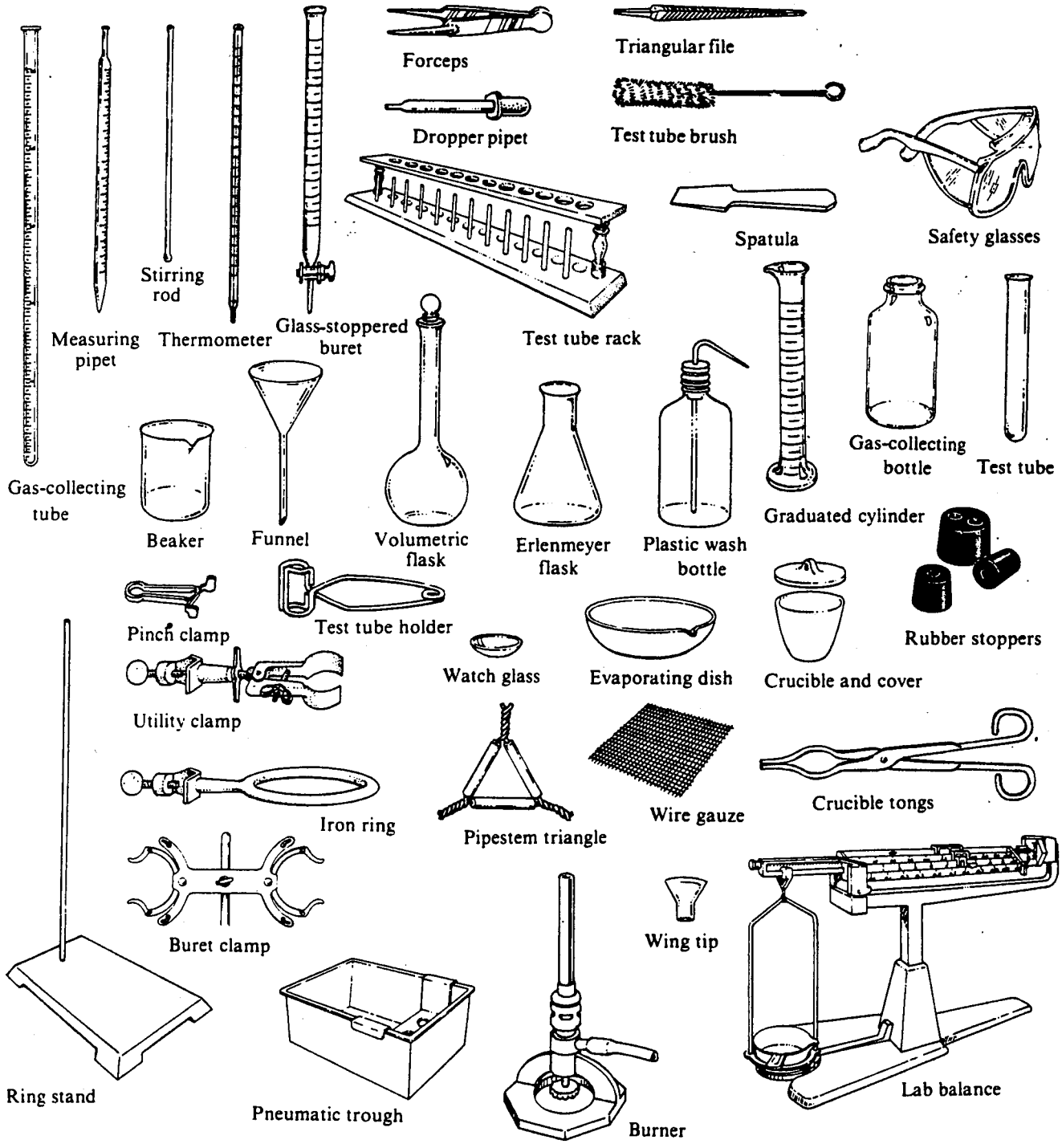
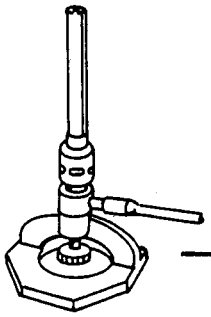


Figure 1. Laboratory equipment.

LABORATORY EQUIPMENT

Name _____

Label the lab equipment below.



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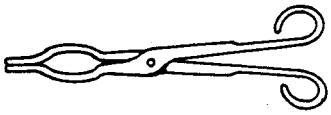

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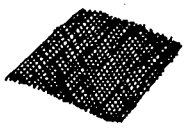

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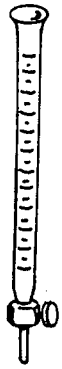

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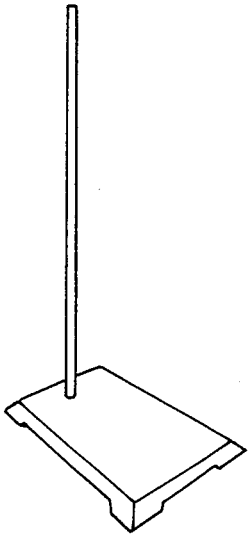

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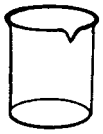

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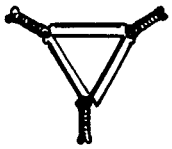

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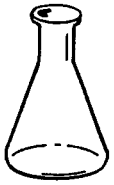

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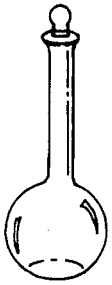

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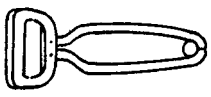

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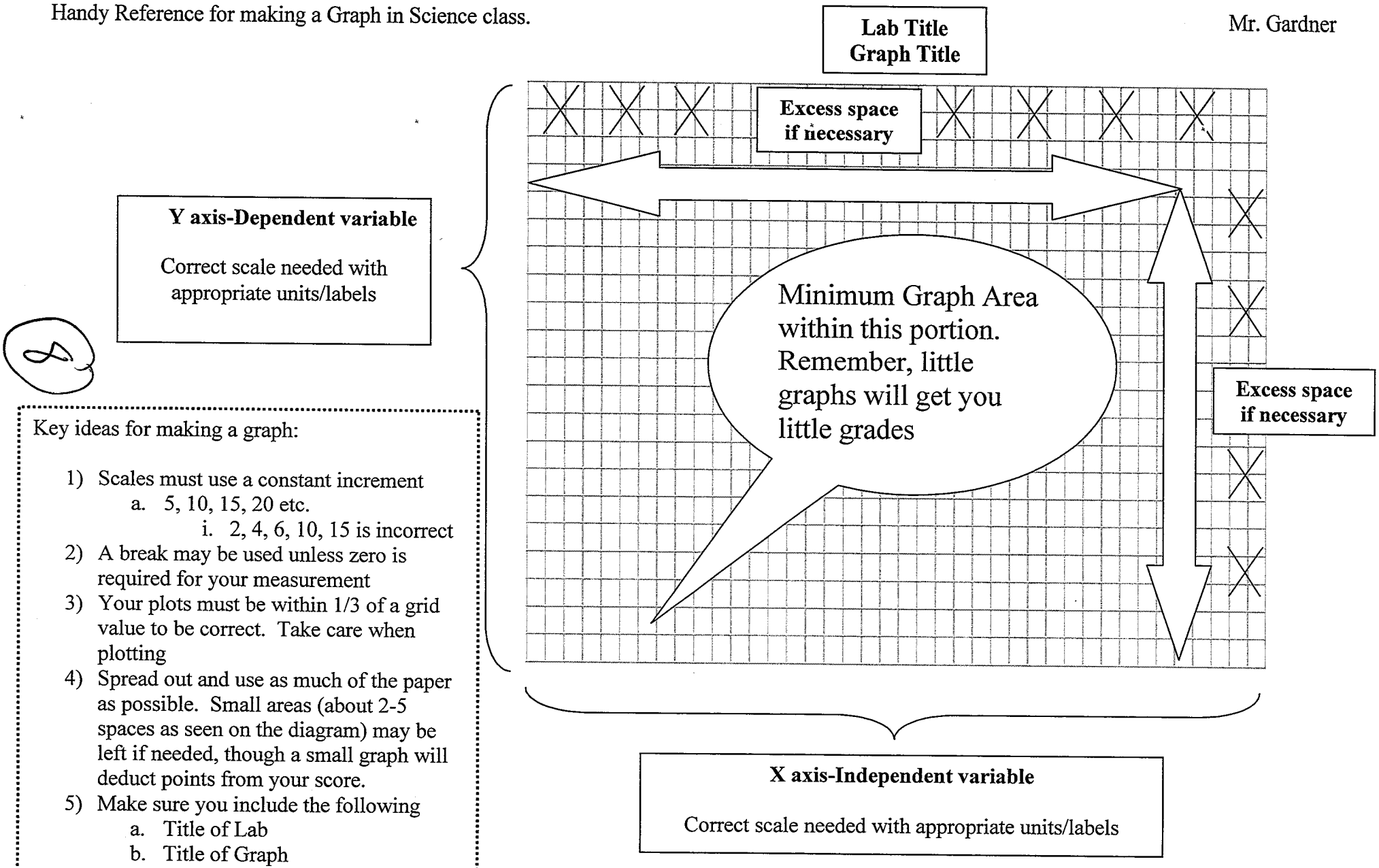
7

Handy Reference for making a Graph in Science class.

Mr. Gardner



Key ideas for making a graph:

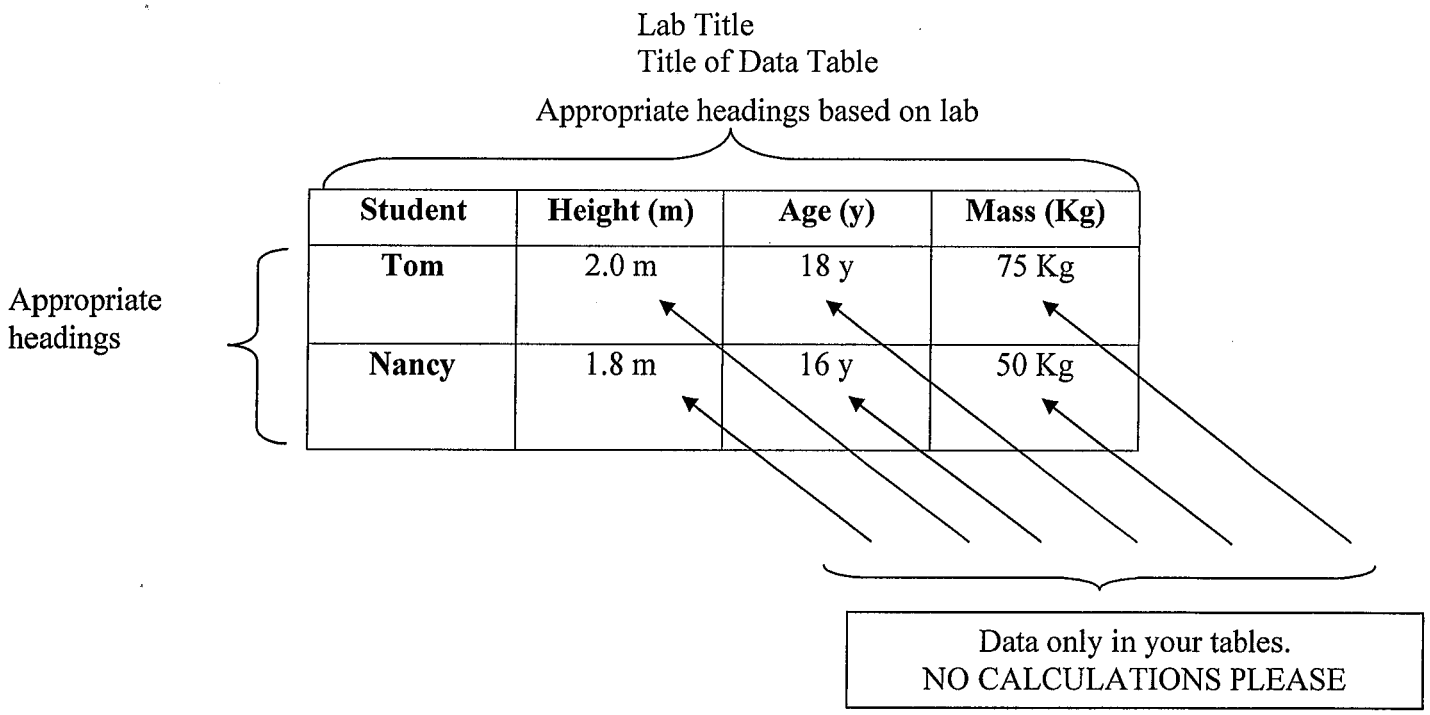
1) Scales must use a constant increment
    a. 5, 10, 15, 20 etc.
        i. 2, 4, 6, 10, 15 is incorrect
2) A break may be used unless zero is required for your measurement
3) Your plots must be within 1/3 of a grid value to be correct. Take care when plotting
4) Spread out and use as much of the paper as possible. Small areas (about 2-5 spaces as seen on the diagram) may be left if needed, though a small graph will deduct points from your score.
5) Make sure you include the following
    a. Title of Lab
    b. Title of Graph
        i. Must be meaningful
            ex. Mass vs. Volume
    c. Axis with appropriate scales

Handy Reference for making a Table in Science Class Mr. Gardner

Directions to Make a table using Microsoft word

1) Click on Table
2) Insert → Table
    a. Select correct number of rows and columns you desire

Many other steps following this will allow for individual designs and style within your tables. Experiment!!!



- Work is not shown within a table, only specific data determined through experimentation.
    - Calculations for this work should be typed/written out and included in the data section
    - Remember to include equations, substitutions and units throughout.

Table title should be meaningful (not just Table 1)

Density

Density is a physical property of matter. Most commonly density refers either to the mass per unit volume (mass density) or the number of objects (e.g., atoms, molecules) per unit volume (number density). We will focus out attention on mass density. The mass density has the units mass/volume. Since volume has the units length "cubed" then the SI unit of mass density is kg/m^3. More common units of density are g/ml or g/l. Substances have differnt densities. In fact the density of a substance can often be used to help idenify it. Below is a table of densities of common materials:

Densities of Some Common Substances	
Substance	Density (g/ml)
ice (0 0C)	0.917
water (4.0 0C)	1.0000
gold	19.31
helium (25 0C)	0.000164
dry air (25 0C)	0.001185
Human fat	0.94
Cork	0.22 - 0.26
table sugar	1.59
balsa wood	0.12
earth	5.54

An important example is water. The above table states that liquid water has a mass of 1 g in every ml. Thus 2 ml of water has a mass of 2 g etc.. Table sugar is more dense than water by about 60 percent. Density does not depend upon size. For example the water in a swimming pool has the same density a glass of that swimming pool water.

Calculations with density are straight forward and involve the formula for density namely D=m/V, where D=density, m= mass and V = volume.

Try the Following for Practice!!!

1. What is the volume of a nugget of gold that has a mass of 3.45 g? The density of gold can be looked upon as a conversion factor from mass to volume i.e.,
2. A light substance is found to weigh 23 g and to have a volume of 0.192 liters. What is the substance?
3. What is the mass of 1 liter of sugar?
4. The density of air at sea level is 1.21 kg m-3. Calculate the mass of air, in kg, contained in a room with dimensions 5 m × 6 m × 2.5 m.

Name_____ Date_____

Mr. Gardner [redacted]

Density of Elements

Atomic Number	Symbol	Name	Density** (g/cm^3)
1	H	hydrogen	0.00009
2	He	helium	0.000179
3	Li	lithium	0.534
4	Be	beryllium	1.8477
5	B	boron	2.340
6	C	carbon	3.513
7	N	nitrogen	0.00125
8	O	oxygen	0.001429
9	F	fluorine	0.001696
10	Ne	neon	0.0009
11	Na	sodium	0.971
12	Mg	magnesium	1.738
13	Al	aluminum	2.698
14	Si	silicon	2.329
15	P	phosphorus	1.820
16	S	sulfur	2.070
17	Cl	chlorine	0.003214
18	Ar	argon	0.001783
19	K	potassium	0.862
20	Ca	calcium	1.550
21	Sc	scandium	2.989
22	Ti	titanium	4.540
23	V	vanadium	6.100
24	Cr	chromium	7.190
25	Mn	manganese	7.440
26	Fe	iron	7.874
27	Co	cobalt	8.900
28	Ni	nickel	8.902
29	Cu	copper	8.960
30	Zn	zinc	7.133
31	Ga	gallium	5.907
32	Ge	germanium	5.323
33	As	arsenic	5.780
34	Se	selenium	4.790
35	Br	bromine	3.122
36	Kr	krypton	0.00375
37	Rb	rubidium	1.532
38	Sr	strontium	2.540
39	Y	yttrium	4.469
40	Zr	zirconium	6.506

Atomic Number	Symbol	Name	Density** (g/cm^3)
41	Nb	niobium	8.570
42	Mo	molybdenum	10.220
43	Tc	technetium	11.500
44	Ru	ruthenium	12.370
45	Rh	rhodium	12.410
46	Pd	palladium	12.020
47	Ag	silver	10.500
48	Cd	cadmium	8.650
49	In	indium	7.310
50	Sn	tin	7.310
51	Sb	antimony	6.691
52	Te	tellurium	6.240
53	I	iodine	4.930
54	Xe	xenon	0.0059
55	Cs	cesium	1.873
56	Ba	barium	3.594
57	La	lanthanum	6.145
Elements 58–71 have been omitted.			
72	Hf	hafnium	13.310
73	Ta	tantalum	16.654
74	W	tungsten	19.300
75	Re	rhenium	21.020
76	Os	osmium	22.590
77	Ir	iridium	22.560
78	Pt	platinum	21.450
79	Au	gold	19.320
80	Hg	mercury	13.546
81	Tl	thallium	11.850
82	Pb	lead	11.350
83	Bi	bismuth	9.747
84	Po	polonium	9.320
85	At	astatine	—
86	Rn	radon	0.00973
87	Fr	francium	—
88	Ra	radium	5.000
89	Ac	actinium	10.060
Elements 90 and above have been omitted.			

*Boiling point at standard pressure

**Density at STP

11

Some common densities

Substance	Density (g/cm^3)
Alcohol	0.8
Aluminum	2.7
Bamboo	0.4
Bone	1.9
Brass	8.4
Brick	1.8
Carbon Tetrachloride	1.6
Copper	8.9
Cork	0.25
Corn oil	0.922
Ethyl alcohol	0.79
Gasoline	0.66-0.69
Glycerine	1.3
Glue	1.27
Gold	19.3
Ice	0.92
Iron	7.8
Kerosene	0.8
Lead	11.3
Mercury	13.6
Milk	1.03
Paraffin	0.9
Porcelain	2.4
Quartz	2.6
Rubber	1.2
Salt	2.2
Sea Water	1.03
Silver	10.5
Steel	7.4
Sugar	1.6
Turpentine	0.9
Water	1
Wood	0.7
*Pine	0.4
*Walnut	0.5
*Cherry	0.5
*Cedar	0.4
*Plywood	0.5
*Balsa	0.2
*Metal Slab	2.4
*Granite	2.3

* approximate values

Name_____ Block_____

Density questions

Directions: Show all work and circle your answer.

1. A piece of silver has a mass of 2800 grams and occupies a volume of 266 cm^3. What is the density of the silver?

2. A flask is pre-weighed and is found to be 15.25 grams. After 5 mL of ethyl alcohol are added, the flask now weighs 19.17 grams. Calculate the density of ethyl alcohol.

3. 200 mL of water are poured from a graduated cylinder into a square container that has the dimensions of 10cm * 5cm * 5cm. Would the water fill up this container?

Name_____ Block_____

Density questions

Directions: Show all work and circle your answer.

1. A piece of silver has a mass of 2800 grams and occupies a volume of 266 cm^3. What is the density of the silver?

REPEAT Questions

2. A flask is pre-weighed and is found to be 15.25 grams. After 5 mL of ethanol are added, the flask now weighs 19.17 grams. Calculate the density of ethanol.

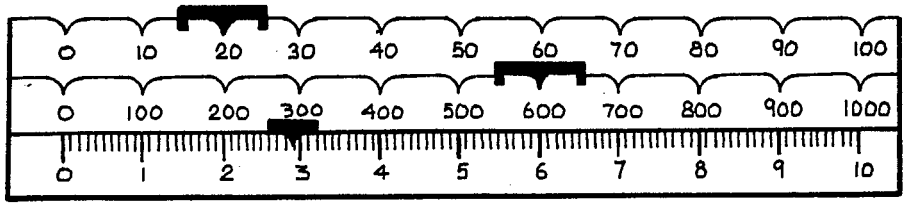
3. 200 mL of water are poured from a graduated cylinder into a square container that has the dimensions of 10cm * 5cm * 5cm. Would the water fill up this container?

THE TRIPLE AND FOUR BEAM BALANCES

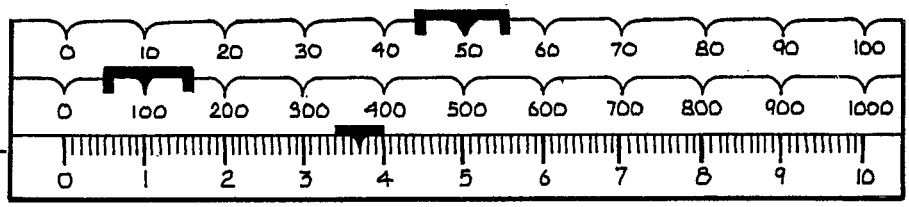
Name _____

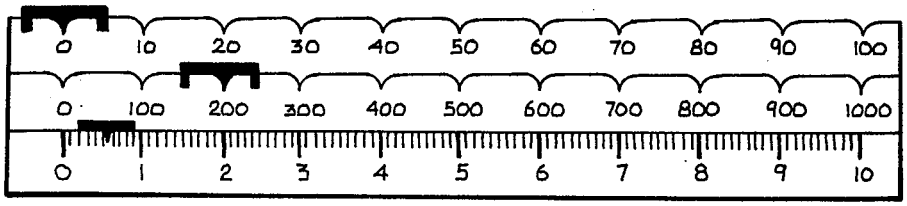
What masses are shown on each of the following balances?

Triple Beam Balance

1) 

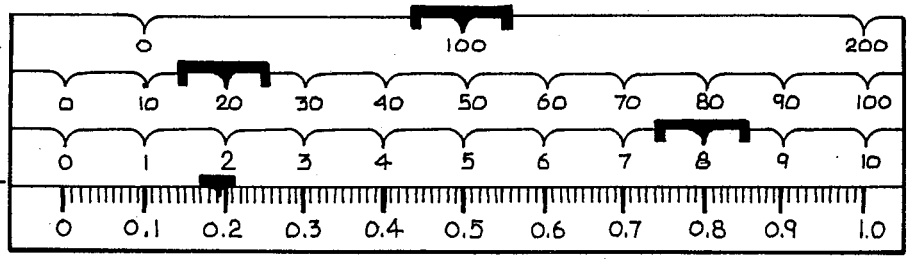
Answer: _____

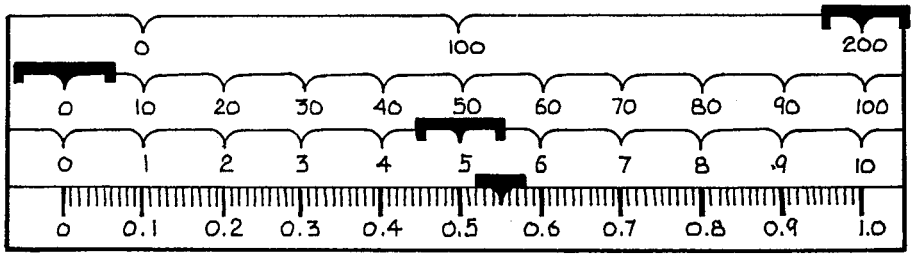
2) Answer: _____ 

3) 

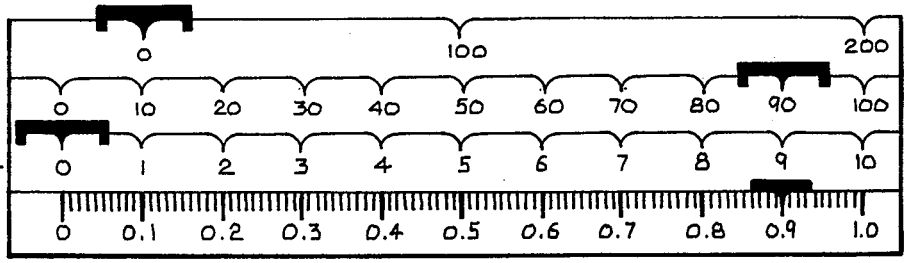
Answer: _____

Four Beam Balance

4) Answer: _____ 

5) 

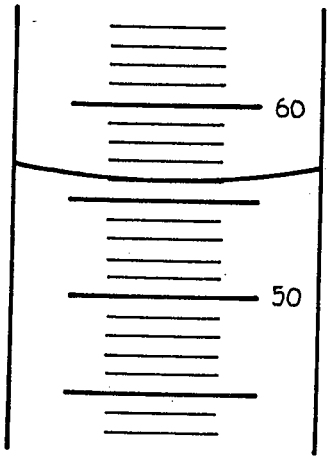
Answer: _____

6) Answer: _____ 

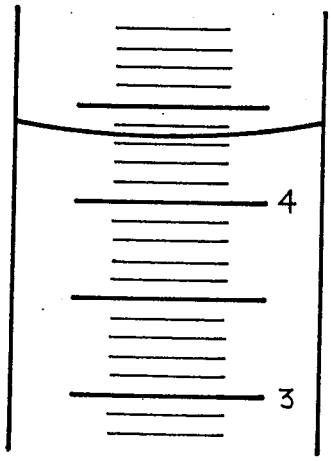
MEASURING LIQUID VOLUME

Name _____

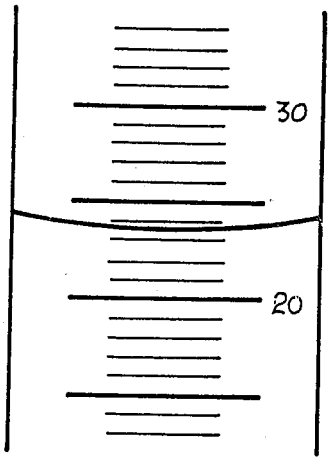
What volume is indicated on each of the graduated cylinders below? The unit of volume is mL.



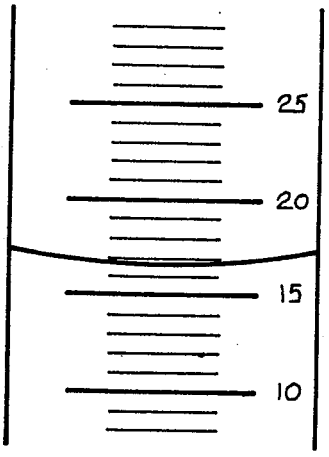
a) _____



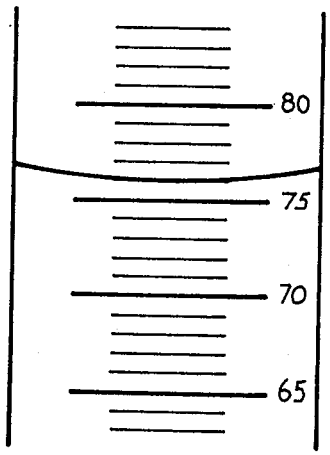
b) _____



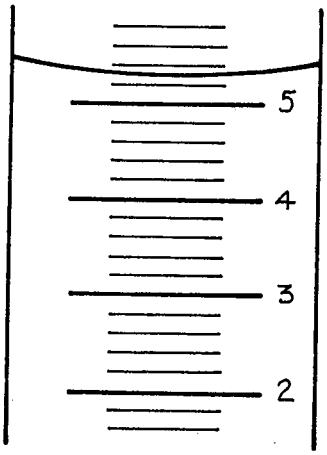
c) _____



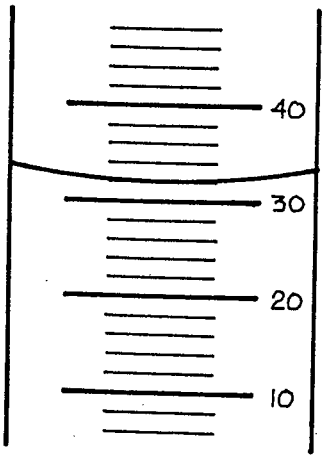
d) _____



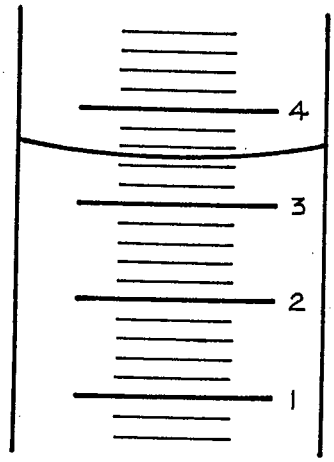
e) _____



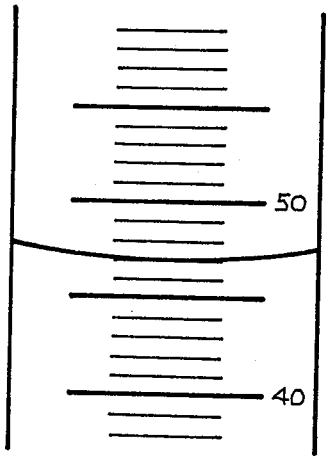
f) _____



g) _____



h) _____

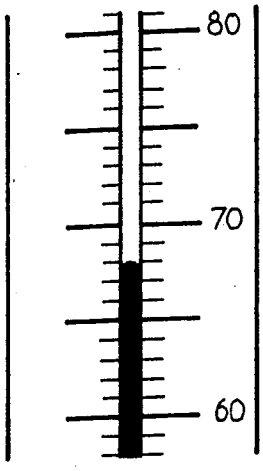


i) _____

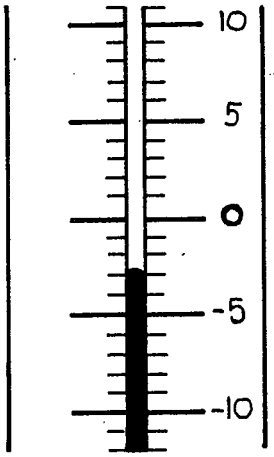
READING THERMOMETERS

Name _____

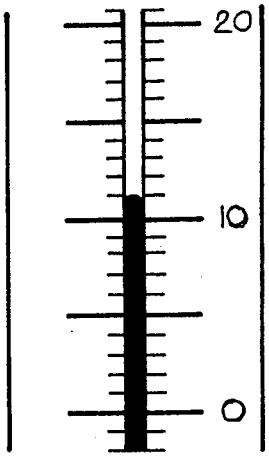
What temperature is indicated on each of the thermometers below?



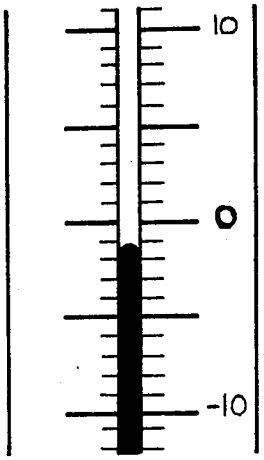
a) _____



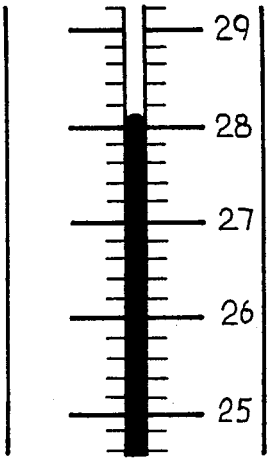
b) _____



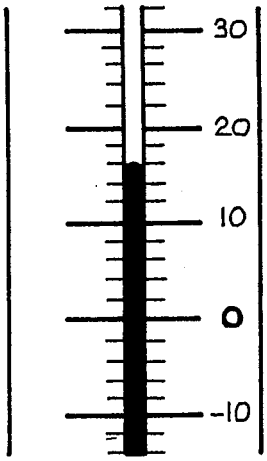
c) _____



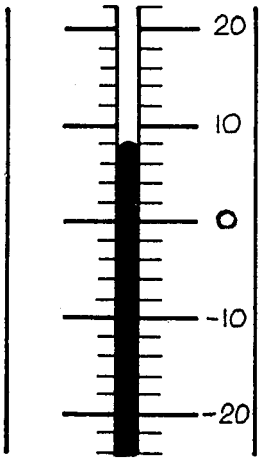
d). _____



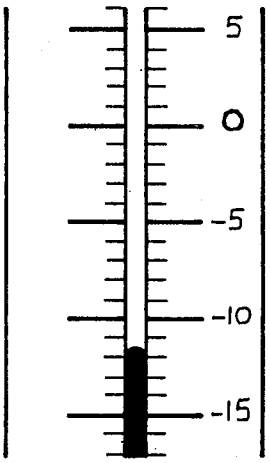
e) _____



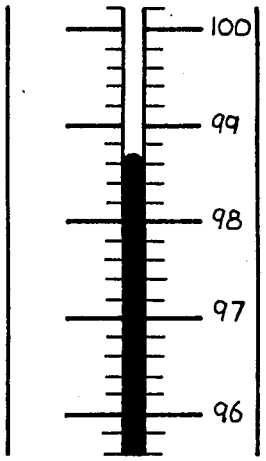
f) _____



g) _____



h) _____



i) _____

Significant Figures/Rounding Rules Name_____

I. All non-zero digits are significant. 125

2. If a zero represents a measured quantity, it is a significant figure.
100. ← **Notice the decimal point marking the final zero**

3. If a zero locates the decimal point, it is **not** a significant figure.
0.0025

100 ← **No decimal point here, so only 1 sig. fig.**

4. ZERO WITHIN A NUMBER. A zero between any of the other digits in a number is a significant figure 102

5. ZERO AT THE BEGINNING OF A NUMBER. Zeros at the first of a number are not significant figures. O.OO205

6. ZERO AT THE END OF A NUMBER. Zeros at the right of the decimal point, at the end of a number, are significant figures. 12.100

***** In the examples by each rule above, the significant figures are underlined*****

Significant Figures with addition (+) and subtraction (-):

Write down the numbers being certain to keep like decimal places under each other. Add or subtract. Note the column which, in reading from left to right, contains the first estimated figure. Round the sum or difference so that this (first estimated figure) column determines the last decimal place of the answer. The final answer is rounded off in this column.

Significant Figures with Multiplication (*) and division (/):

The number of significant figures in the product or quotient obtained from measured quantities is the same as the number of significant figures in the quantity having the smallest number of significant figures.

17

Significant Figures Name _____

Significant figures (also known as significant digits) are the digits that are measured with certainty plus one digit that is estimated.

Rules to Determine the Number of Significant Figures

1) Digits from 1 to 9 are always significant

ex: 345 3 significant figures
1234 4 significant figures

2) Zeros between 2 non-zero digits are significant.

ex: 4.004 4 significant figures
1203 4 significant figures

3) Zeros at the end of a number are only significant if there is a decimal point present.

Ex: 300 1 significant figure
300. 3 significant figures
300.00 5 significant figures

4) For numbers less than one, leading zeros are not significant.

Ex: 0.005 1 significant figure
0.10 2 significant figures

How many significant figures are in the following numbers?

(1) **400**
(2) **200.0**
(3) **0.001**
(4) **218**
(5) **320**
(6) **0.00530**
(7) **22568**
(8) **22000**
(9) **4755.50**
(10) **91.0020**
(11) **0.092010**
(12) **0.00341**
(13) **1.0040**
(14) **0.00005**
(15) **65000**
(16) **5201**
(17) **40300**
(18) **200300**
(19) **0.4004**
(20) **6000**
(21) **1.00030**
(22) **400.**
(23) **635.000**
(24) **81**

(18)

Significant Figure Calculations

Perform the following calculations and round the answers according to the rules for addition, subtraction, multiplication, and division of significant figures. Show all work.

1) 357.89 + 0.002 =

2) 17.95 + 32.42 + 50 =

3) 5.5 + 3.7 + 2.97 =

4) 84.675 – 3 =

3) 75 – 2.55 =

6) 10 – 9.9 =

7) 50.0 x 2.00 =

8) 2.3 x 3.45 x 7.42 =

9) 1.0007 x 0.009 =

10) 51 / 7 =

11) 208 / 9.0=

12) 0.003 / 5 =

Solve the following calculations: **Show all formulas and work even if a calculator is used!!!**

Give the number of significant figures in the space provided:

1. 402 m _____	2. 0.00402 L _____	3. 0.48 g _____
4. 34.20 mg _____	5. 3200 g _____	6. 0.03 mL _____
7. 0.300 mg _____	8. 480 kg _____	9. 0.0205 g _____

Multiply or divide and show the final answer with the appropriate number of significant figures in the space provided. Use a calculator if necessary.

10. 17 x 204 _____	11. 1.7 x 4924 _____	12. 0.005 x 8888 _____
13. 0.03 / 0.0050 _____	14. 0.074 / 692 _____	15. 0.50000 / 250 _____

Add or subtract the following and show the appropriate number of significant figures in your final answer in the space provided.

16. 3.40g
0.022 g
+0.5 g
Answer:_____

17. 102.45 g
2.44 g
+1.9999 g
Answer:_____

18. 42.306 mL
-1.22 mL
Answer:_____

Complete the following with the use of a calculator, as needed, and place answers in space provided.

19. A rectangular solid measures 13.4 cm by 11.0 cm by 2.2 cm. Calculate the volume of the solid.

Answer:_____

20. If the density of liquid mercury is 13.6 g/mL, what is the mass in grams of 3426 mL of the liquid?

Answer:_____

21. A copper cylinder, 12.0 cm in radius, is 44.0 cm long. If the density of copper is 8.90 g/ cm3, calculate the mass in grams of the cylinder.

Answer:_____

20

Scientific Notation Worksheet I

Write the following in Scientific Notation:

1. 190000
2. 528
3. 10
4. 4400000
5. 9700
6. 0.1
7. 49
8. 0.056
9. 0.001
10. 0.00000035
11. 0.00077
12. 0.00000512
13. 0.01
14. 123
15. 0.0001
16. 0.113

Write the following in Standard Notation:

17. 1.986×10^{5}
18. 1.75×10^{7}
19. 2.5×10^{1}
20. 1.411×10^{-3}
21. 1×10^{3}
22. 1.986×10^{8}
23. 3.33×10^{8}
24. 9.28×10^{-8}
25. 6.275×10^{-6}
26. 1×10^{0}

Write the following in <u>Proper</u> Scientific Notation:

27. 2101×10^{-16}
28. $0.0000594 \times 10^{-16}$
29. 3.86×10^{-22}
30. 305.1×10^{7}
31. $0.00000827 \times 10^{19}$
32. 2511×10^{12}

Scientific Notation Calculations I

Multiplication:

1. $(3 \times 10^{5})(3 \times 10^{6})$

2. $(4 \times 10^{-6})(4 \times 10^{-4})$

Division:

3. $\dfrac{(3.45 \times 10^{8})}{(6.74 \times 10^{-2})}$

4. $\dfrac{(4.7 \times 10^{-2})}{(5.7 \times 10^{-6})}$

Addition:

5. $8.41 \times 10^{3} + 9.71 \times 10^{4}$

6. $8.2 \times 10^{2} + 4.0 \times 10^{3}$

Subtraction:

7. $5.11 \times 10^{2} - 4.2 \times 10^{2}$

8. $6.3 \times 10^{-2} - 2.1 \times 10^{-1}$

CHEMISTRY WORKSHEET Name_____

FACTOR-LABEL PROBLEMS

Instructions: Make each of the factor-label conversions in the spaces provided. Show your setup in a **single line**. Show all **conversions** and **cancellations** in this line. Report your answer to the proper number of **significant figures** in the answer spaces provided for each problem.

Use the following equivalents:

12 in = 1 ft	3 ft = 1 yd	5280 ft = 1 mi	32 oz = 1 qt (liq)
4 qt = 1 gal	2 pt = 1 qt	454 g = 1 Lb	946 mL = 1 qt
2.54 cm = 1 in	1 mL = 1 cm^3	1 Lb = 16 oz (dry)	1 mL = 1 g (water)

1.) How many inches are represented by 0.042 meters?

Ans._____

2.) What is the mass in kilograms of a 3400 Lb car?

Ans._____

3.) How many gallons are represented by 0.056 liters?

Ans._____

4.) How far in meters is 2.00 miles?

Ans._____

5.) How many gallons of water would it take to fill a cube that has a volume of 64,000,000 cm^3?

Ans._____

6.) The distance from Syracuse to Buffalo is 153 miles. Calculate the distance in km.

Ans._____

7.) How many ounces are in a gallon of milk?

Ans._____

8.) The distance from the North Pole to the equator is 1,000,000 m. What would be the distance in feet?

Ans._____

Answers:

1.) 1.7 in
2.) 1,500 kg
3.) 0.015 gal
4.) 3,220 m
5.) 16,900 gal
6.) 246 km
7.) 128 oz
8.) 3,300,000 ft

Note: These are rounded-off answers. If yours round off to the same thing, you are correct.

Name_____ FACTOR LABEL PRACTICE

1) The average student is in class 330 minutes/day.

A) How many hours is this?
B) How many seconds is this?

2) How many seconds are there in 1 year?

3) Chicago uses 1.2×10^9 gallons of water/day. How many gallons per second must be pumped from the lake every second to supply the city?

4) Change 60 miles/ hour to _____ft/sec (clue: both miles and hour must be change to other units)

5) How old is Mr. Gardner in seconds? (Born on June 28, 1975)

25

Name_____ Mr. Gardner

Problem Solving Factor Label Method

1. Calculate the number of minutes in the entire month of May

2. A sign in a town gives the speed limit at 50 km/hr. What is this speed in centimeters per second?

3. A chemistry instructor provides each student with 8 test tubes at the beginning of the school year. If there are 28 students per class, how many test tubes are required for three chemistry classes?

4. Near a lake on an old building a sign reads, "Rowboats for rent, $1.75 per half-hour." What will it cost to rent a rowboat for five hours?

5. In your favorite restaurant, a sandwich you like costs $1.25. If you order two sandwiches, how many quarters must you pay? How many dimes?

6. What is the cost in dollars for the nails used to build a fence 125 meters long if it requires 20 nails per meter? Assume that 40 nails are sold per box at a cost of 75 cents per box.

7. At a meeting, 28 people are each given 3 pens. If there are 8 pens in one package, priced at $1.88 per package, what is the total cost of giving away pens?

8. An object is traveling at a speed of 7500 centimeters per second. Convert the value to kilometers per minute.

20

Name_____ Chem Quiz

Factor Label Method: Complete the following, showing all steps and conversions

1) How many seconds are in a leap year?

2) If you get pulled over in front of St. Anne's for doing 50mi/hr, what is that in km/min?
Remember 5280 ft in a mile and 1inch is equal to 2.54 cm.

3) 45cm=_____km

4) 13kg=_____mg

Scientific Method: List the following in correct order from left to right in the spaces below.

A) Form hypothesis B) Make observations C) Define problem
D) Experiment E) Draw conclusions

_____ _____ _____ _____ _____

Density: Solve the following. Hint→ Use Table S in your periodic table

1) What is the density of a material that has a mass of 48 grams and displaces 12ml of water?

2) What is the volume of iron present if I have a piece with a mass of 200.0g?

Vocabulary: Give an example for each of the following (words and diagrams may be used)

1) Homogenous

2) Heterogenous

3) Precise

27

Measurement-Factor label **Name**_____
Mr. Gardner

_____**1.** A sample of water is being heated from 20°C to 30°C, and the temperature is recorded every 2 minutes. Which table would be most appropriate for recording the data?

Time (min)	Temp (°C)
0	
2	
4	
6	
8	
10	

(1)

Time (min)	Temp (°C)
20	
22	
24	
26	
28	
30	

(2)

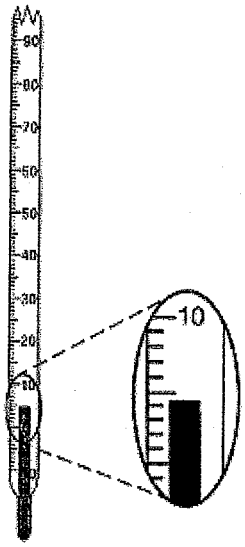
Temp (°C)	Time (min)
0	
2	
4	
6	
8	
10	

(3)

Temp (°C)	Time (min)
20	
22	
24	
26	
28	
30	

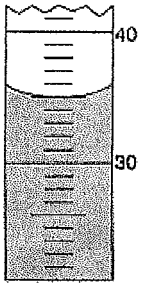
(4)

_____**2.** The accompanying diagram represents a Celsius thermometer recording a certain temperature. What is the correct reading of the thermometer?



1. 5°C
2. 4.3°C
3. 0.3°C
4. 4°C

_____**3.** The diagram shown represents a portion of a 100-milliliter graduated cylinder. What is the reading of the meniscus?



1. 35.0 mL
2. 36.0 mL
3. 44.0 mL
4. 45.0 mL

28

Solve the following problems with the use of your calculator and reference tables where necessary.

SHOW ALL WORK AND REMEMBER YOUR UNITS!!!

1 inch = 2.54 cm

3 feet = 1 yard

1. If your heat beats 45 times a minute how many times will it beat in an hour?

2. You just bought a new 60 inch plasma flat screen TV. What size would the screen be in meters?

3. If a pizza delivery driver travels 196.8 kilometers a day how many centimeters will he drive in a week?

4. Imagine that water is leaking from a container, at a rate of 1.2 ml/hour. If this rate does not change, how many liters of water will be lost in a week?

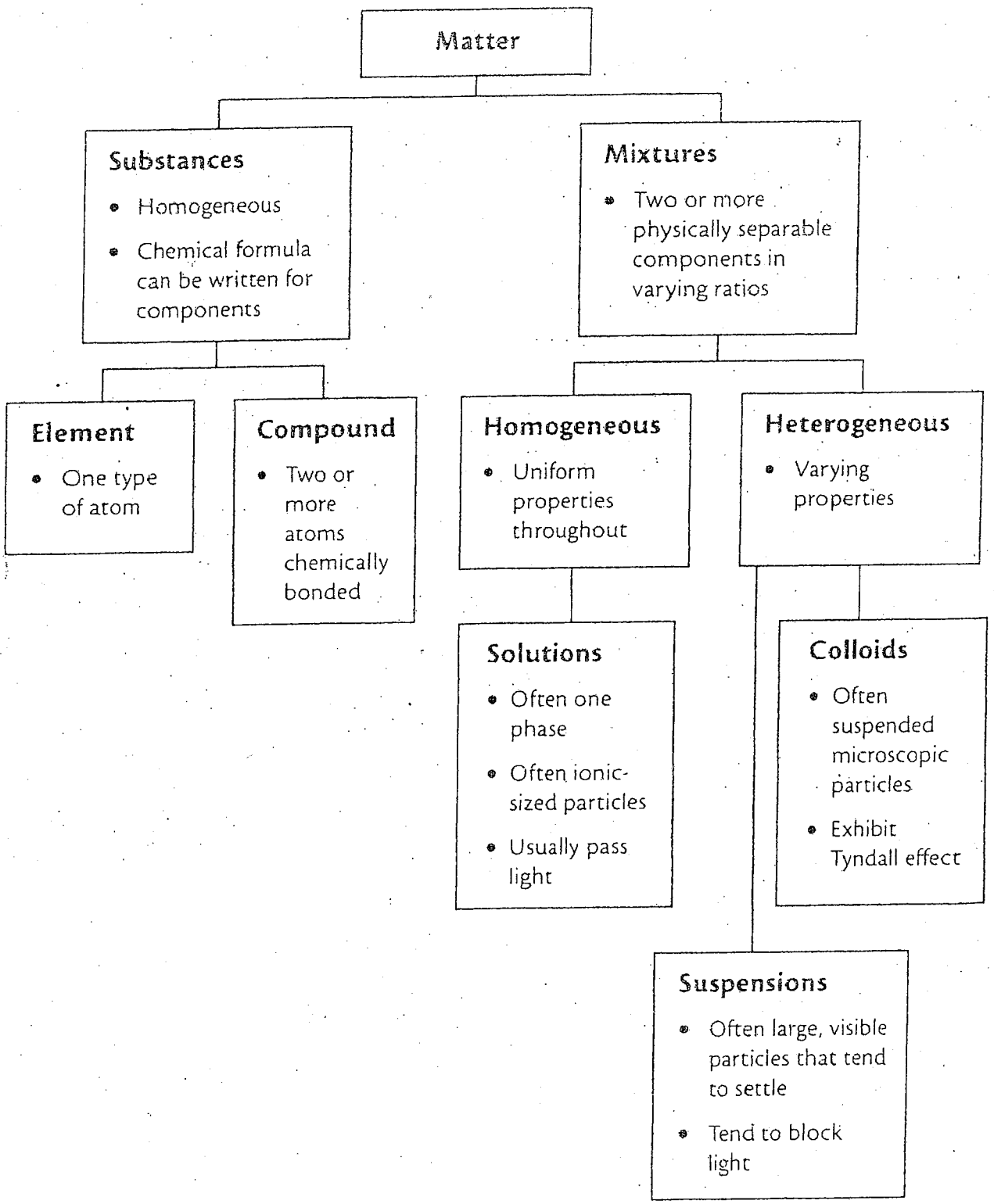
Circle the best choice:

qualitative/ quantitative: information that cannot be counted or measured

qualitative/quantitative: information that can be either counted or measured

Unit 2

17 Types of Matter



Properties of Matter

Property	Description	Example
Electrical conductivity	ability to carry electricity	Copper is a good electrical conductor, so it is used in wiring.
Heat conductivity	ability to transfer energy as heat	Aluminum is a good heat conductor, so it is used to make pots and pans.
Density	mass-to-volume ratio of a substance; measure of how tightly matter is "packed"	Lead is a very dense material, so it is used to make sinkers for fishing line.
Melting point	temperature at which a solid changes state to become a liquid	Ice melts to liquid water at the melting point of water.
Boiling point	temperature at which a liquid boils and changes state to become a gas at a given pressure	Liquid water becomes water vapor at the boiling point of water.
Index of refraction	extent to which a given material bends light passing through it	The index of refraction of water tells you how much light slows and bends as it passes through water.
Malleability	ability to be hammered or beaten into thin sheets	Silver is quite malleable, so it is used to make jewelry.
Ductility	ability to be drawn into a thin wire	Tantalum is a ductile metal, so it is used to make fine dental tools.

Mixture Definitions

Mixture- The combination of two or more substances that can be separated by physical means (distillation, filtration, etc.). Each substance keeps its own properties. Example: kool-aid, salt water.

Heterogeneous Mixture- A mixture that has a non-uniform composition. Ex: Concrete, chocolate chip ice cream.

Suspension- A heterogeneous mixture with particles large enough to settle out and be filtered. Ex: mud.

Colloid – A heterogeneous mixture with small particles present that don't settle out. Colloids scatter light with the Tyndall effect. Example: jello, fog.

Homogeneous Mixture – A mixture with uniform composition throughout. The substance dissolves completely. A homogeneous solution does not show the Tyndall effect.

Solutions are homogeneous mixtures.

Solute – The dissolved substance in a solution.

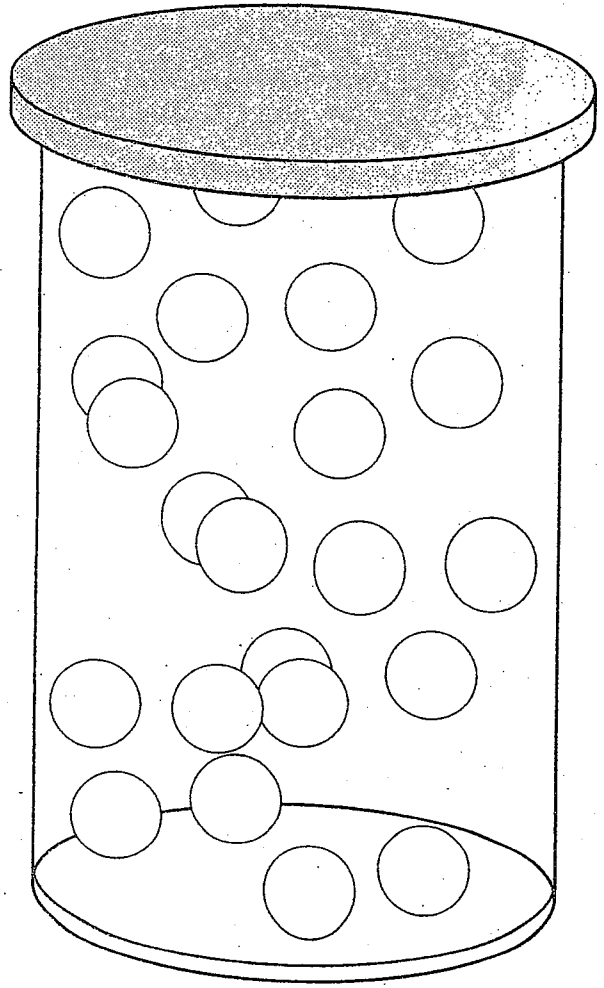
Solvent – The substance that does the dissolving in the solution.

Aqueous Solution – A solution that has water as the solvent.

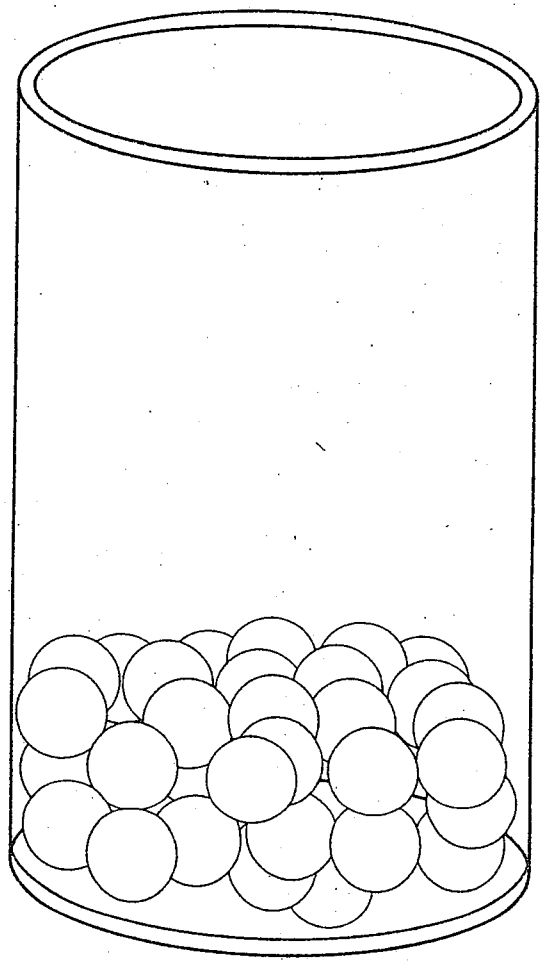
35 STATES OF MATTER

Particles of a gas travel in a completely random manner; particles of a liquid appear to vibrate around moving points; and particles of a solid appear to vibrate around fixed points.

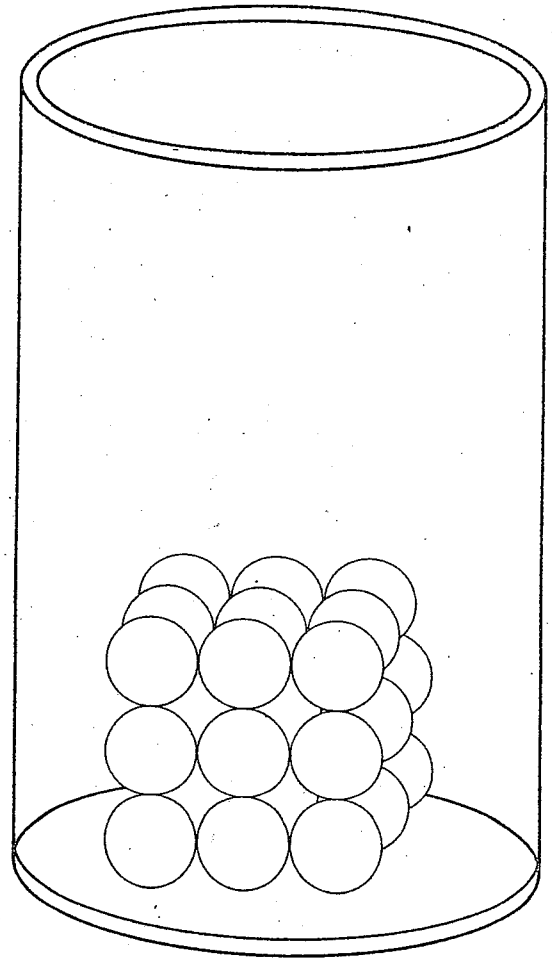
33



Gas



Liquid



Solid

Name_____ Specific Heat and Calorimetry wkst

1. Calculate the final temperature when 50 mL of water at 40°C are added to 25 mL of water at 20°C. (Hint: if you assume a closed system where the Law of Conservation of Energy is obeyed, then, the heat lost by the "hot" water will be equal to the heat gained by the "cold" water to achieve thermal equilibrium)

OMIT #1

2. A piece of metal weighing 51.0 g at a temperature of 48.6°C, was placed in a calorimeter containing 20.00 mL of water at 22.1°C. The final equilibrium temperature was found to be 28.2°C. What is the specific heat of the metal?

3. How much heat is needed to raise the temperature of 20.0g of water from 5.0°C→20.0°C?

4) How much heat is released by 200.0g of solid aluminum as it cools from 200.0°C→150.0°C? (c= 0.897J/g·°C for aluminum)

5) What is the specific heat of copper if 385J were applied to a 5g sample to raise its temperature from 0.0°C→ 200.0°C

34

Name: _____

Worksheet: Heat of Fusion and Heat of Vaporization

1) How much heat energy, in joules, is needed to melt 6.0 grams of solid water (ice) at its melting point?

2) How much heat energy, in joules, is absorbed when 20 grams of solid ethanol is converted to liquid at its melting point? (The heat of fusion of solid ethanol is 104 J/g)

3) What is the total number of joules of heat energy needed to change 150 grams of ice to water at 0° C?

4) How many joules of heat energy are required to change 40 grams of ice at 0° C to water at the same temperature?

5) 30.139 kJ are added to a large block of ice at 0° C. How much of the ice will melt?

35

Review Questions

Write and balance equations for the following synthesis reactions.

48. hydrogen and bromine forming hydrogen bromide

49. fluorine and argon forming argon trifluoride

50. sulfur and oxygen forming sulfur dioxide

51. calcium and chlorine forming calcium chloride

52. nickel and oxygen forming nickel(II) oxide

Write and balance equations for the following decomposition reactions.

53. decomposition of water into hydrogen and oxygen

54. decomposition of aluminum oxide into aluminum and oxygen

55. decomposition of sodium chloride into sodium and chlorine

56. decomposition of ammonia into hydrogen and nitrogen

57. decomposition of mercury(II) oxide into mercury and oxygen

For each of the following, indicate whether or not a single replacement reaction will occur.

58. aluminum and hydrochloric acid

59. silver and magnesium chloride

60. chromium and lead(II) nitrate

61. silver and gold(III) chloride

62. chlorine and sodium iodide

For the following, write a balanced chemical equation to show how the ions would combine in a double replacement equation.

63. sodium bromide and silver nitrate form sodium nitrate and silver bromide

64. potassium carbonate and calcium nitrate form potassium nitrate and calcium carbonate

65. ammonium sulfate and barium chloride form ammonium chloride and barium sulfate

66. barium nitrate and potassium chromate form barium chromate and potassium nitrate

67. sodium hydroxide and calcium chloride form sodium chloride and calcium hydroxide

68. Identify each of the following as unbalanced equations for a synthesis (*S*), decomposition (*D*), single replacement (*SR*), or double replacement (*DR*) reactions.

(a) $Zn + HCl \rightarrow ZnCl_2 + H_2$
(b) $NaClO_3 \rightarrow NaCl + O_2$
(c) $P_4 + Cl_2 \rightarrow PCl_3$
(d) $HCl + Mg(OH)_2 \rightarrow MgCl_2 + H_2O$
(e) $BaO + SO_3 \rightarrow BaSO_4$
(f) $Pb + AgNO_3 \rightarrow Ag + Pb(NO_3)_2$
(g) $AgNO_3 + Na_2CrO_4 \rightarrow Ag_2CrO_4 + NaNO_3$
(h) $Al + Fe_3O_4 \rightarrow Al_2O_3 + Fe$
(i) $NO_2 \rightarrow 2NO + O_2$
(j) $NaN_3 \rightarrow Na + N_2$
(k) $Pb(NO_3)_2 + KI \rightarrow KNO_3 + PbI_2$
(l) $CaO + CO_2 \rightarrow CaCO_3$
(m) $MgCO_3 \rightarrow MgO + CO_2$
(n) $Na + Cl_2 \rightarrow NaCl$
(o) $HNO_3 + Mg(OH)_2 \rightarrow Mg(NO_3)_2 + H_2O$
(p) $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$
(q) $Fe + O_2 \rightarrow Fe_3O_4$
(r) $Cl_2 + KI \rightarrow KCl + I_2$
(s) $Ba(NO_3)_2 + Na_2SO_4 \rightarrow BaSO_4 + NaNO_3$
(t) $Ag_2O \rightarrow Ag + O_2$

Review Questions

1. Which substance has a definite shape and a definite volume at STP? (1) NaCl(*aq*) (2) $Cl_2(g)$ (3) $CCl_4(\ell)$ (4) $AlCl_3(s)$

2. At STP, which element has a definite shape and volume? (1) Ag (2) Hg (3) Ne (4) Xe

3. Which sample is most likely to take the shape of and occupy the total volume of its container? (1) $CO_2(g)$ (2) $CO_2(\ell)$ (3) $CO_2(aq)$ (4) $CO_2(s)$

4. Which substance takes the shape of and fills the volume of any container into which it is placed? (1) $H_2O(\ell)$ (2) $CO_2(g)$ (3) $I_2(s)$ (4) $Hg(\ell)$

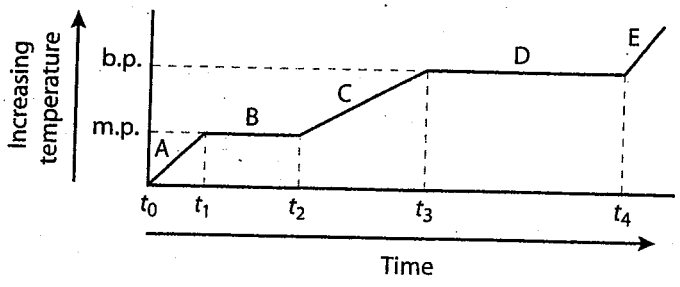
5. As a substance changes from a liquid to a gas, the average distance between molecules (1) decreases (2) increases (3) remains the same

6. In which phase are the particles the most random? (1) solid (2) liquid (3) gas

7. Which phase change represents sublimation? (1) $H_2O(\ell) \rightarrow H_2O(s)$ (2) $H_2O(\ell) \rightarrow H_2O(g)$ (3) $I_2(s) \rightarrow I_2(g)$ (4) $I_2(s) \rightarrow I_2(\ell)$

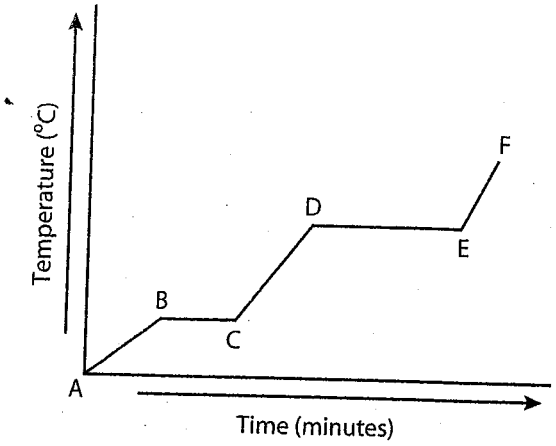
8. Which phase change represents sublimation? (1) $NH_3(\ell) \rightarrow NH_3(g)$ (2) $CO_2(s) \rightarrow CO_2(g)$ (3) $KI(s) \rightarrow KI(\ell)$ (4) $H_2O(\ell) \rightarrow H_2O(s)$

9. A solid substance initially at a temperature below its melting point is heated at a constant rate. The heating curve for the substance is shown in the graph below.



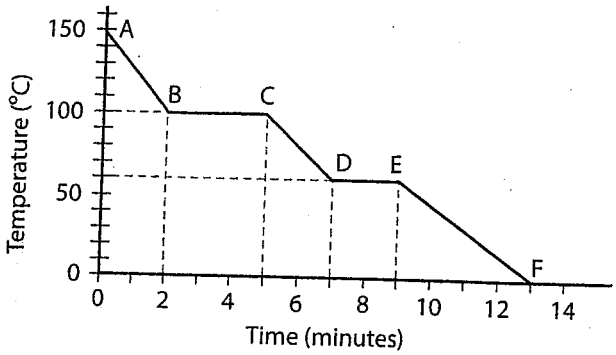
Which portions of the graph represent times when heat is absorbed and potential energy increases while kinetic energy remains constant? (1) A and B (2) B and D (3) A and C (4) C and D

10. A solid substance initially at a temperature below its melting point is heated at a constant rate. The heating curve for the substance is shown in the graph on the following page.



Which segment of the graph represents a time when both the solid and liquid phases are present? (1) AB (2) BC (3) DE (4) EF

11. A gaseous substance initially at a temperature above its boiling point is cooled at a constant rate. The cooling curve for the substance is shown below.



How much time passes between the first appearance of the liquid phase of the substance and the presence of the substance completely in its solid phase? (1) 5 minutes (2) 2 minutes (3) 7 minutes (4) 4 minutes

12. The heat of fusion is defined as the energy required (at constant temperature) to change a (1) gas to a liquid (2) gas to a solid (3) solid to a gas (4) solid to a liquid

13. As ice cools from 273 K to 263 K, the average kinetic energy of its molecules (1) decreases (2) increases (3) remains the same

14. Which occurs as a substance melts? (1) It changes from a solid to a liquid and heat is absorbed. (2) It changes from a solid to a liquid and heat is released. (3) It changes from a liquid to a solid and heat is absorbed. (4) It changes from a liquid to a solid and heat is released.

15. Which phase change is endothermic? (1) gas → solid (2) gas → liquid (3) liquid → solid (4) liquid → gas

16. Which phase change is exothermic? (1) $H_2O(s) \rightarrow H_2O(\ell)$ (2) $H_2O(\ell) \rightarrow H_2O(s)$ (3) $H_2O(s) \rightarrow H_2O(g)$ (4) $H_2O(\ell) \rightarrow H_2O(g)$

Review Questions

17. Which is not a form of energy? (1) light (2) temperature (3) electricity (4) heat

18. Which unit is used to express the amount of energy absorbed or released during a chemical reaction? (1) degree (2) torr (3) gram (4) joule

19. Which term represents a form of energy? (1) heat (2) degree (3) kilojoule (4) temperature

20. The minimum number of fixed reference points required to establish the Celsius temperature scale for a thermometer is (1) 1 (2) 2 (3) 3 (4) 4

21. What are the fixed reference points on the Celsius thermometer? (1) 32 and 100 (2) 0 and 212 (3) 32 and 212 (4) 0 and 100

22. The difference between the boiling point and the freezing point of pure water at standard pressure is (1) 32 K (2) 273 K (3) 100 K (4) 373 K

23. What is the freezing point of water on the Kelvin scale at standard pressure? (1) 0 K (2) 32 K (3) 100 K (4) 273 K

24. When the temperature of an object changes by 100°C, the same temperature change in Kelvins would be (1) 100 K (2) 173 K (3) 273 K (4) 373 K

25. Energy is added to a substance. Compared to the Celsius temperature of the substance, the Kelvin temperature (1) will always be 273 greater (2) will always be 273 lower (3) will have the same reading at 0 (4) will have the same reading at 273

26. What Kelvin temperature is equal to −73°C? (1) 100 K (2) 173 K (3) 200 K (4) 346 K

27. Which temperature is equal to 20 K? (1) −253°C (2) −293°C (3) 253°C (4) 293°C

28. Different masses of copper and iron have the same temperature. Compared to the average kinetic energy of the copper atoms, the average kinetic energy of the iron atoms is (1) less (2) greater (3) the same

29. The average kinetic energy of water molecules increases when (1) $H_2O(s)$ changes to $H_2O(\ell)$ at 0°C (2) $H_2O(\ell)$ changes to $H_2O(s)$ at 0°C (3) $H_2O(\ell)$ at 10°C changes to $H_2O(\ell)$ at 20°C (4) $H_2O(\ell)$ at 20°C changes to $H_2O(s)$ at 10°C

Questions for Regents Practice

Part A

1. Which set of properties does a substance such as $CO_2(g)$ have?

(1) definite shape and definite volume

(2) definite shape but no definite volume

(3) no definite shape but definite volume

(4) no definite shape and no definite volume

2. A liquid is poured from a volumetric flask into a beaker. Which of the following is true?

(1) It retains its original volume and shape.

(2) It retains its original volume, but its shape changes.

(3) It retains its original shape, but its volume changes.

(4) Both the volume and shape change.

3. The heat required to change 1 gram of a solid at its normal melting point to a liquid at the same temperature is called the heat of

(1) vaporization (3) reaction

(2) fusion (4) formation

4. Which statement best describes the molecules of H_2O in the solid phase?

(1) They move slowly in straight lines.

(2) They move rapidly in straight lines.

(3) They are arranged in a regular geometric pattern.

(4) They are arranged in a random pattern.

5. As the temperature of a substance rises, the average kinetic energy of the particles making up the substance

(1) increases (3) remains the same

(2) decreases

6. When a substance melts, it undergoes a process known as

(1) condensation (3) sublimation

(2) fusion (4) vaporization

7. Which phase change is accompanied by the release of heat?

(1) $H_2O(s) \rightarrow H_2O(\ell)$ (3) $H_2O(s) \rightarrow H_2O(g)$

(2) $H_2O(\ell) \rightarrow H_2O(s)$ (4) $H_2O(\ell) \rightarrow H_2O(g)$

8. Which of the following is a unit of heat?

(1) torr (3) gram

(2) degree (4) joule

9. Which of the following behave most like ideal gases?

(1) oxygen and hydrogen (3) oxygen and nitrogen

(2) helium and hydrogen (4) helium and nitrogen

Part B

10. Which is the equivalent of 750. joules?

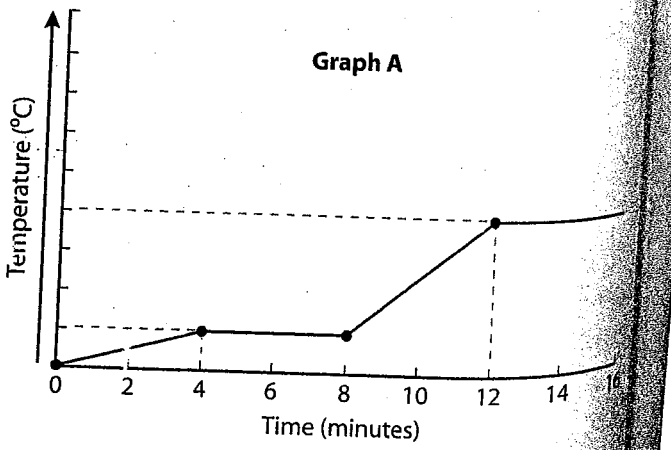
(1) 0.750 kJ (3) 7.50 kJ

(2) 75 kJ (4) 750. kJ

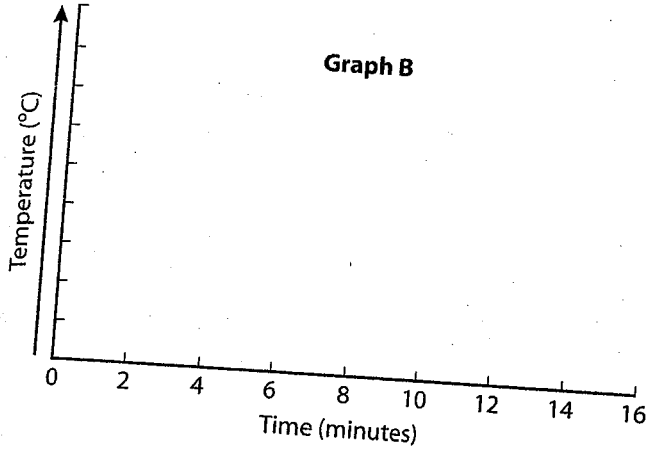
39

Questions 22–24 are based on the following graphs.

Graph A



Graph B



22. Graph A shows the temperature change of 10.0 g of a substance from below its melting point until after it begins boiling when heat is added at a rate of 200. J/min. Use Graph B to plot the change in temperature for the same substance when heat is added at a rate of 400. J/min. [2]

23. Based on Graph A, what is the heat of fusion of the substance? [1]

24. How do the heats of fusion in Graph A and Graph B compare? [1]

26. What data would a student need to collect in an experiment to determine the specific heat capacity of a substance? [1]

11. As a solid is heated at a constant rate, its temperature increases from 10°C to 25°C, remains at 25°C for 5 minutes, and then increases to beyond 45°C. Based on this information, what conclusion can be drawn about the substance?

(1) Its melting point is 45°C.

(2) Its boiling point is 45°C.

(3) Its melting point is 25°C.

(4) Its boiling point is 25°C.

12. A liquid's freezing point is −38°C and its boiling point is 357°C. How many Kelvins are there between the boiling point and the freezing point of the liquid?

(1) 319

(2) 395

(3) 592

(4) 668

13. Which Celsius temperature is equivalent to 323 K?

(1) 50°C

(2) 212°C

(3) 273°C

(4) 596°C

14. When steam condenses to water, the surrounding temperature

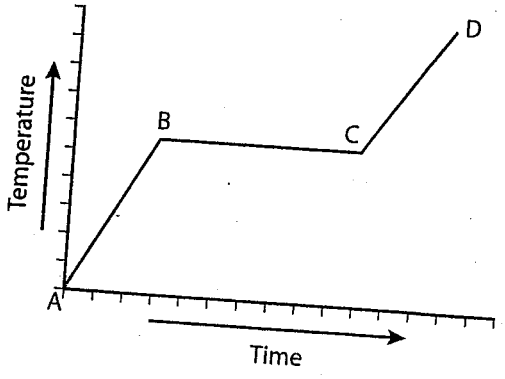
(1) decreases

(2) increases

(3) remains the same

Show all work for the following questions. Use appropriate units and follow operations with significant figures.

15. The following graph represents the relationship between temperature and time as heat was added uniformly to a substance. At the beginning, the substance was a solid below its melting point. Describe the changes in kinetic and potential energy of the substance during time periods AB and BC.



16. How many joules of energy are needed to convert 50.0 g of water at 50°C to 65°C?

40

Name _____ Date _____ Class _____

1 Chapter Test

Chemistry and You

Multiple Choice

On the line at the left, write the letter of the answer that best completes each sentence.

_____ 1. A tentative answer to a scientific question is called a
- **a.** theory.
- **b.** conclusion.
- **c.** law.
- **d.** hypothesis.

_____ 2. Experimentation involves the testing of
- **a.** laws.
- **b.** variables.
- **c.** classification.
- **d.** conversions.

_____ 3. The most important rule in lab safety is to
- **a.** wear gloves.
- **b.** know your equipment.
- **c.** follow instructions.
- **d.** clean up when you're finished.

_____ 4. An example of an SI base unit is a
- **a.** kilogram.
- **b.** force.
- **c.** pressure.
- **d.** power.

_____ 5. An example of a derived quantity is
- **a.** area.
- **b.** mass.
- **c.** length.
- **d.** time.

_____ 6. Measurements are always uncertain because
- **a.** instruments aren't precise.
- **b.** measurements involve some estimation.
- **c.** people don't know how to use instruments.
- **d.** both a and b

_____ 7. A measurement that is very close to the standard accepted value is said to be
- **a.** analogous.
- **b.** precise.
- **c.** accurate.
- **d.** uncertain.

_____ 8. Zeros are always considered significant digits when they
- **a.** function as place keepers.
- **b.** occur between two nonzero numbers.
- **c.** precede a decimal point.
- **d.** both a and b

Name _____ Date _____ Class _____

1 Chapter Test *(continued)*

_____ **9.** The unit of measure for density is
- **a.** g/s.
- **b.** g/C°.
- **c.** g/cm^3.
- **d.** g/Kg.

_____ **10.** How many significant digits are in the measurement 0.005 030 g?
- **a.** 3
- **b.** 4
- **c.** 5
- **d.** 7

_____ **11.** The technique of dimension analysis is used to
- **a.** convert a very large number into an easier-to-use form.
- **b.** determine the independent variable on a graph.
- **c.** convert units of measurement.
- **d.** develop a testable hypothesis.

True or False

If the statement is true, write "true." If it is false, change the underlined word or words to make the statement true.

_____ **12.** The safety symbol showing a flame should remind you to secure all loose <u>equipment</u>.

_____ **13.** The <u>meter</u> is the SI base unit for measuring length.

_____ **14.** The metric prefix <u>*centi*-</u> means one thousandth, or 0.001.

_____ **15.** A <u>sensitive</u> measuring instrument will give consistent results under the same conditions.

Problems

Express each of the following in scientific notation.

16. 0.000 000 015 cm

17. 1,360,000,000 g

Perform the following calculations, then round off each answer to the proper number of significant digits.

18. 6.15 m x 3.026 m x 0.018 m

19. $\frac{1.6305\ \text{g}}{2.0\ \text{m}^3}$

Name _____ Date _____ Class _____

1 Chapter Test *(continued)*

20. Identify how many significant digits are in each of the following measurements and write your answer on the line.

_____ **a.** 187.032 g

_____ **b.** 0.0601 m^3

_____ **c.** 1.30×10^{-12} kg

_____ **d.** 620 L

Solve each of the following problems as directed. Show all your work.

21. Convert 160.57 g into each of the following units. Use scientific notation where convenient.

a. _____ kg

b. _____ μg

c. _____ mg

22. The density of iron is 7.86 g/cm^3. You are given an unknown metal that has a volume of 30.1 cm^3. What would the mass of the sample be if the metal is iron?

23. Calculate the number of feet in a 5-km race, given the following unit equalities: 1 in = 2.54 cm; 12 in = 1 ft.

Essays

Write your answers to the following questions on a separate sheet of paper.

24. Explain how you would use the scientific method to determine whether or not dogs can see color. Be sure to include each of the steps of the scientific method as discussed in the chapter.

25. Could you generate a natural law from the results you would obtain in question 24? Explain.

Name _____ Date _____ Class _____

Additional Problems

Solve each of the following problems as directed. Show all your work.

26. In medieval Britain, property was often measured in units such as fardells, nookes, yards, and hides. Using the following conversions, calculate the number of fardells in eight hides.

two fardells = one nooke
four nookes = one yard
four yards = one hide

27. If air is 21 percent oxygen by volume and oxygen has a density of 1.31 g/L, what is the volume of a room that contains 80.0 kg of oxygen?

28. If 6.02×10^{23} water molecules have a mass of 18 g, calculate how many water molecules fill a swimming pool that measures 5.0 m x 5.0 m x 5.0 m. The density of water is $\frac{1000\ \text{kg}}{\text{m}^3}$.

29. The volume of a small piece of gold (density = 19.3 g/cm^3) is 24.8 cm^3. What is the volume of a piece of silver (density = 10.5 g/cm^3) that has the same mass as the gold?

Additional Essay

Write your answer to the following question on a separate sheet of paper.

30. Explain why it is not possible for a scientist to state that a scientific theory or law is absolutely and forever correct.

Name _____ Date _____ Class _____

2 Chapter Test

Energy and Matter

Multiple Choice

On the line at the left, write the letter of the answer that best completes each statement.

_____ 1. Energy is defined as the capacity to
- **a.** cause chemical change.
- **b.** exert force.
- **c.** do work.
- **d.** resist gravity.

_____ 2. Which of the following is an example of potential energy?
- **a.** a dry-cell battery in your camera
- **b.** the water behind a dam
- **c.** the gasoline in your car tank
- **d.** all of the above

_____ 3. The SI scale used to measure temperature is the
- **a.** Celsius scale.
- **b.** Fahrenheit scale.
- **c.** Kelvin scale.
- **d.** Calorie scale.

_____ 4. Absolute zero corresponds to
- **a.** 0°C.
- **b.** 273 K.
- **c.** –273 K.
- **d.** 0 K.

_____ 5. The law of conservation of matter states that
- **a.** matter is saved in the form of particles.
- **b.** matter is conserved by chemical processes.
- **c.** matter is very unlike energy.
- **d.** matter is neither created nor destroyed in any process.

_____ 6. The new properties observed during a change of state are not signs of a chemical change because
- **a.** the new properties are temporary.
- **b.** the chemical identity of the substance has not been altered.
- **c.** the color of the substance has changed.
- **d.** a gas is not produced.

_____ 7. Which of the following is an example of a chemical change?
- **a.** steam condensing on the bathroom mirror
- **b.** cracking open an egg and removing the yolk
- **c.** burning a piece of toast
- **d.** water evaporating from a puddle on a hot day

2 *Chapter Test (continued)*

_____ 8. Elements and compounds are both considered pure substances because they
- **a.** cannot be broken down into simpler substances.
- **b.** appear on the periodic table.
- **c.** contain two or more elements combined in a fixed proportion.
- **d.** have a unique set of chemical and physical properties.

_____ 9. Electrolysis could be used to
- **a.** separate water into oxygen and hydrogen gas.
- **b.** break zinc into simpler substances.
- **c.** remove salt from ocean water.
- **d.** both a and b

_____ 10. Both homogeneous and heterogeneous mixtures
- **a.** are blends of two or more pure substances.
- **b.** can be separated by filtration.
- **c.** appear to be made of a single kind of matter.
- **d.** are less common than pure elemental substances.

_____ 11. All of the following are heterogeneous mixtures except
- **a.** air.
- **b.** milk.
- **c.** paint.
- **d.** concrete.

Matching

Match the following elements to their corresponding element symbols. Write the letter on the line.

_____	**12.** potassium	**a.**	Na
_____	**13.** sodium	**b.**	Hg
_____	**14.** copper	**c.**	Pb
_____	**15.** gold	**d.**	K
_____	**16.** iron	**e.**	Cu
_____	**17.** mercury	**f.**	Sn
_____	**18.** tin	**g.**	Fe
_____	**19.** lead	**h.**	Au

Problems

Solve each of the following problems as directed. Show all your work.

20. A runner burns about 10 kilocalories per minute. If the runner completes a race in one hour and fourteen minutes, how many kilocalories did he burn? How many grams of pasta would have provided him with this much energy? (The energy value of pasta is 17,000 J/g.)

46

21. Fill in the chart by supplying each missing temperature conversion.

	°C	K
a. Boiling point of water	100	
b. Freezing point of water		273
c. Room temperature	20	
d. Normal body temperature	37	
e. Absolute zero		0

22. An electric current is passed through a 68.3-g sample of water to separate it into its component parts. If 7.59 g of hydrogen gas is released, how much oxygen is produced?

Short Answer

Answer the following question in the space provided.

23. Describe the steps you would take in order to separate a mixture of water, salt, sand, and paper clips.

_____

_____

_____

Essays

Write your answers to the following questions on a separate sheet of paper.

24. Write a short paragraph describing an ordinary activity that involves at least five physical changes and three chemical changes. Underline the physical changes and circle the chemical changes.

25. Describe three ways in which energy and matter are related.

47

Additional Questions

Answer each of the following questions in the space provided.

26. If burning 1 mL of gasoline releases 1.03×10^{-4} calories, how many joules of potential chemical energy are in one gallon of gasoline? (1 gallon = 3.744 L)

27. What temperature, measured in Fahrenheit degrees, is exactly twice the measurement that the temperature would be on the Celsius scale? ($°F = \frac{9}{5}°C + 32$)

28. A swimming pool with the dimensions 12 m x 5 m x 2 m is filled with water. If the water temperature is 20°C, how many calories of energy are needed to raise the temperature to 25°C? (The density of water is 1 kg/m^3.)

29. Describe how distillation can be used to separate solid impurities from water and retain the pure water.

_____

_____

_____

30. Wine contains alcohol, but foods cooked in a wine sauce usually do not. How can you account for this apparent contradiction to the law of conservation of matter?

_____

_____

_____

Additional Essay

Write your answer to the following question on a separate sheet of paper.

31. Define absolute zero and explain why it has not been reached.