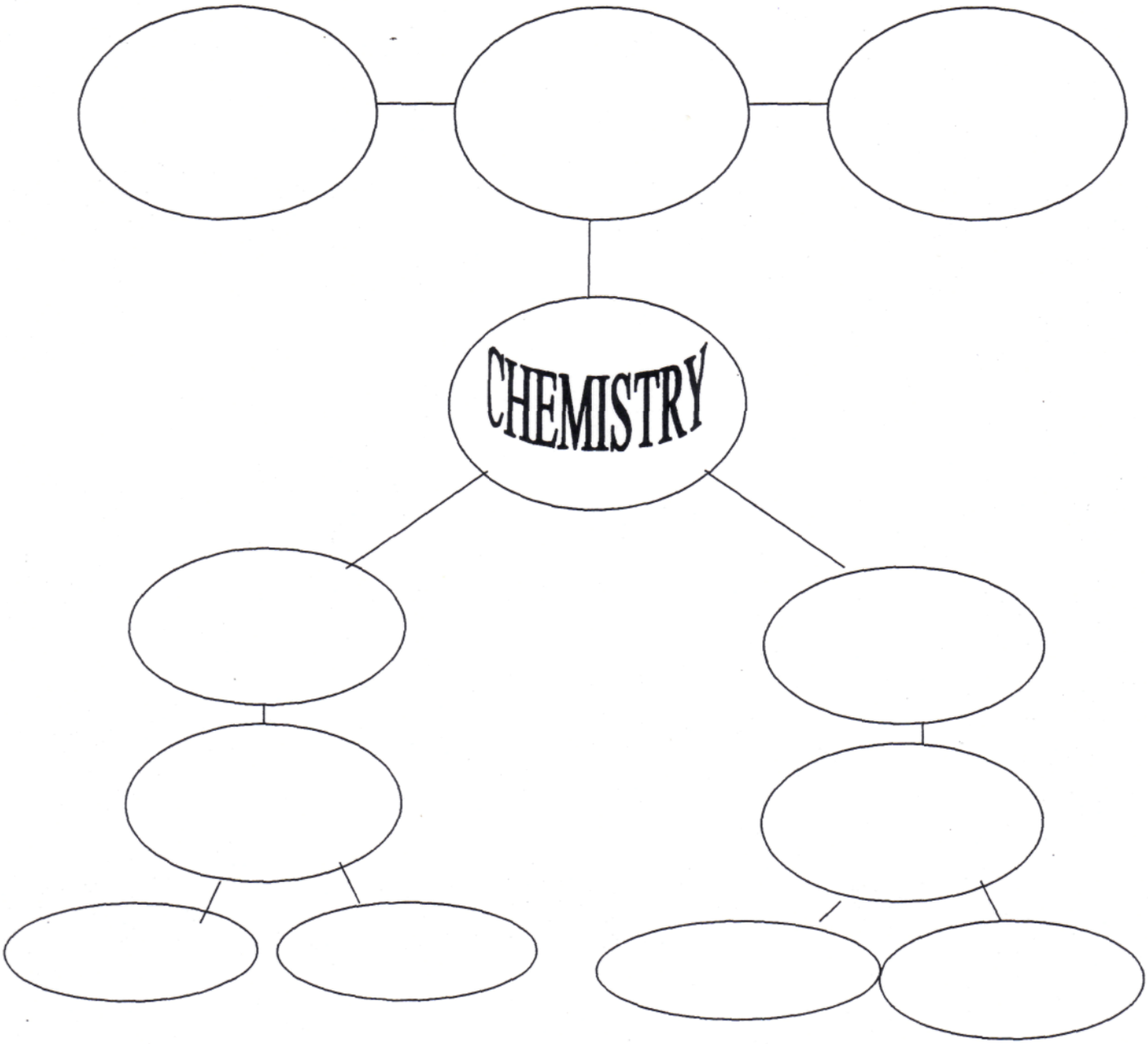
THE MEANING OF CHEMISTRY

Do now: When you hear the word chemistry what do you think of?

Definition of the word CHEMISTRY:

Activity: As a class we are going to complete the following concept map.



Name: _____

Date: _____

Closure activity for the MEANING OF CHEMISTRY

Directions: Using the glossary in your textbook (or a dictionary if the words are unavailable) define the following terms:

1. Chemistry:

2. Physical science:

3. Science:

4. Technology:

5. Matter:

6. Inertia:

7. Mass:

8. Energy:

9. Endothermic:

10. Exothermic:

PHYSICAL VS. CHEMICAL CHANGES

Name _____

In a physical change, the original substance still exists, it has only changed in form. In a chemical change, a new substance is produced. Energy changes always accompany chemical changes.

Classify the following as being a physical or chemical change.

1. Sodium hydroxide dissolves in water. _____
2. Hydrochloric acid reacts with potassium hydroxide to produce a salt, water and heat. _____
3. A pellet of sodium is sliced in two. _____
4. Water is heated and changed to steam. _____
5. Potassium chlorate decomposes to potassium chloride and oxygen gas. _____
6. Iron rusts. _____
7. When placed in H_2O, a sodium pellet catches on fire as hydrogen gas is liberated and sodium hydroxide forms. _____
8. Evaporation _____
9. Ice melting _____
10. Milk sours. _____
11. Sugar dissolves in water. _____
12. Wood rotting _____
13. Pancakes cooking on a griddle _____
14. Grass growing in a lawn _____
15. A tire is inflated with air. _____
16. Food is digested in the stomach. _____
17. Water is absorbed by a paper towel. _____

Name _____ Date _____ Class _____

CHAPTER 4 REVIEW ACTIVITY Text Reference: Section 4-7

Properties

Recall that *physical properties* can be observed without producing new substances. *Chemical properties* describe how a substance interacts (or fails to interact) with other substances to produce new substances. *Extensive properties* depend upon the amount of matter in the sample; *intensive properties* do not.

Classify each of properties listed below as *extensive physical, intensive physical,* or *chemical.*

1. Color — 1. _____
2. Combustibility — 2. _____
3. Hardness — 3. _____
4. Density — 4. _____
5. Mass — 5. _____
6. Melting point — 6. _____
7. Ductility — 7. _____
8. Volume — 8. _____
9. Failure to react with other substances — 9. _____
10. Odor — 10. _____
11. Weight — 11. _____
12. Malleability — 12. _____
13. Tendency to corrode — 13. _____

Some of the measured properties of a given substance are listed below. Write the general name describing each property. Select the names from the properties listed for Exercises 1–13 above.

14. 15 dm^3 — 14. _____
15. Can easily be hammered into sheets. — 15. _____
16. 2.8 g/cm^3 — 16. _____
17. Burns when heated in the presence of O_2. — 17. _____
18. Stinks when heated. — 18. _____
19. Can be scratched by a diamond. — 19. _____
20. 500°C — 20. _____
21. Can easily be drawn into a wire. — 21. _____

PHYSICAL VS. CHEMICAL PROPERTIES

Name _____

A physical property is observed with the senses and can be determined without destroyin the object. For example, color, shape, mass, length and odor are all examples of physicc properties.

A chemical property indicates how a substance reacts with something else. The original substance is fundamentally changed in observing a chemical property. For example, the ability of iron to rust is a chemical property. The iron has reacted with oxygen, and the original iron metal is changed. It now exists as iron oxide, a different substance.

Classify the following properties as either chemical or physical by putting a check in the appropriate column.

	Physical Property	Chemical Property
1. blue color		
2. density		
3. flammability		
4. solubility		
5. reacts with acid to form H_2		
6. supports combustion		
7. sour taste		
8. melting point		
9. reacts with water to form a gas		
10. reacts with a base to form water		
11. hardness		
12. boiling point		
13. can neutralize a base		
14. luster		
15. odor		

Name _____ Date _____ Class _____

CHAPTER 4 Text Reference: Section 4-5

Non-SI Supplementary Problems

The density of a liquid is often expressed as the number of grams per milliliter of the substance (g/mL). Because the density of a gas is often much less than that of a liquid, the density of a gas is usually expressed as grams per liter (g/L).

Example: A sample of an unknown liquid is found to have a volume of 87.3 mL and a mass of 99.1 grams. Calculate its density, expressing your answer in terms of the units given.

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

$$= \frac{99.1\ \text{g}}{87.3\ \text{mL}}$$

$$= 1.135\ \text{g/mL} = 1.14\ \text{g/mL}$$

Exercises

Solve the following problems on a separate sheet of paper, showing all work. Express your answers in the correct units with the appropriate number of significant figures.

1. What is the density of an element if a sample having a mass of 43.2 g has a volume of 96.5 mL?

2. A sample of a gas has a volume of 4.0 L and a mass of 4.922 g. What is its density?

3. Mercury has a density of 13.6 g/mL. What is the volume of a sample of mercury that has a mass of 2242 g?

4. If a liquid has a density of 0.880 g/cm^3, what volume of this liquid would have a mass of 54 g?

5. What is the mass of 84 mL of a liquid if its density is 1.25 g/mL?

6. What is the mass of 25 mL of oxygen gas if its density is 1.43 g/L?

7. A student determines the mass and volume of three samples of a liquid to be:

Sample	Volume	Mass
A	116 mL	85 g
B	168 mL	101 g
C	158 mL	115 g

Could all of these be samples of the same substance? If not, which could be?

8. A gas is confined in a rectangular tank 25.0 cm long, 8.0 cm high and 10.4 cm wide.

If the density of the gas is 19.3 g/L, what is the mass of the gas?

9. The density of an acid is 1.85 g/mL. What volume of the acid would have a mass of 64 g?

10. If 40.0 mL of a liquid with a mass of 44.8 g was mixed with 50.0 mL of a liquid having a mass of 48.0 g, what would the density of the resulting liquid be?

11. A student measures the mass and volume of samples of three liquids to be:

Liquid	Volume	Mass
A	48.5 mL	37.2 g
B	12.8 mL	174.1 g
C	64.7 mL	71.2 g

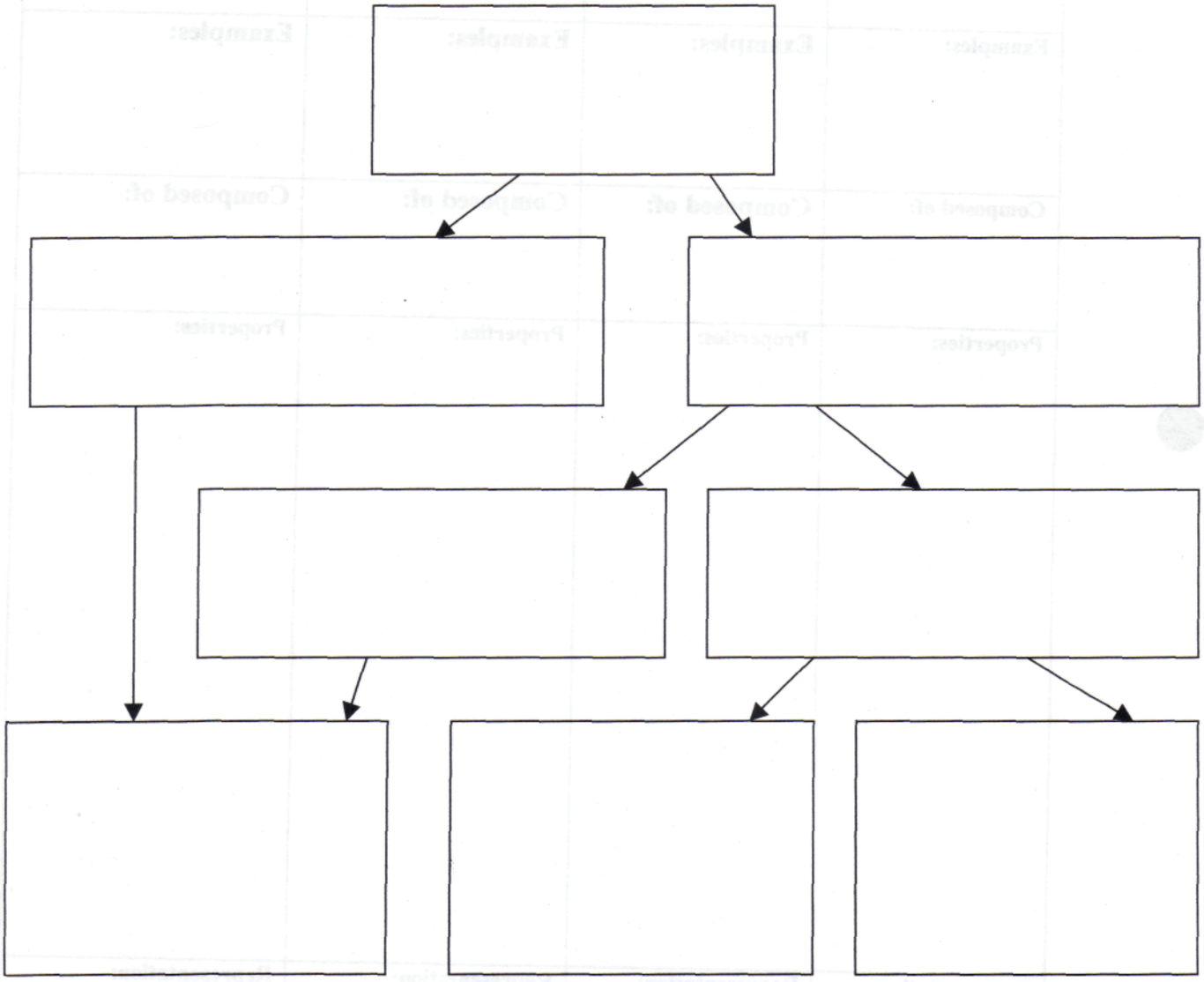
a. What is the density of each liquid?

b. These liquids do not mix. If the samples are poured into one container, which will rise to the top? Which will sink to the bottom?

CLASSIFYING MATTER

Use the terms below to complete the following diagram that illustrates the classification of matter.

Matter	Mixtures	Heterogeneous Matter	Pure Substances
Compounds	Elements	Solutions	Homogeneous matter



CLASSES OF MATTER

ELEMENTS	COMPOUNDS	HOMOGENEOUS MIXTURES	HETEROGENEOUS MIXTURE
Definition:	**Definition:**	**Definition:**	**Definition:**
Examples:	**Examples:**	**Examples:**	**Examples:**
Composed of:	**Composed of:**	**Composed of:**	**Composed of:**
Properties:	**Properties:**	**Properties:**	**Properties:**
Representation:	**Representation:**	**Representation:**	**Representation:**

Classification of Matter

Use words from the list to fill in the blanks in the paragraphs.

Word List

chemical property	heterogeneous matter	physical property
compound	homogeneous matter	property
element	intensive property	substance
extensive property	mixture	

Matter that has uniform characteristics throughout is called ___(1)___. Matter that has parts with different characteristics is called ___(2)___. A characteristic by which a variety of matter is recognized is called a(n) ___(3)___. A characteristic that depends upon the amount of matter in the sample is called a(n) ___(4)___. A characteristic that does not depend upon the amount of matter is called a(n) ___(5)___. A characteristic that can be observed without producing new kinds of matter is called a(n) ___(6)___. A characteristic that depends on how a kind of matter changes its composition (or fails to change its composition) during interactions with other kinds of matter is called a(n) ___(7)___.

Matter can also be classified according to the basic types of matter it contains. A simple substance that cannot be broken down into other substances by chemical means is called a(n) ___(8)___. A chemical combination of simple substances is called ___(9)___. A physical combination of different substances that retain their individual properties is called a(n) ___(10)___. Either an element or a compound may be referred to as a(n) ___(11)___.

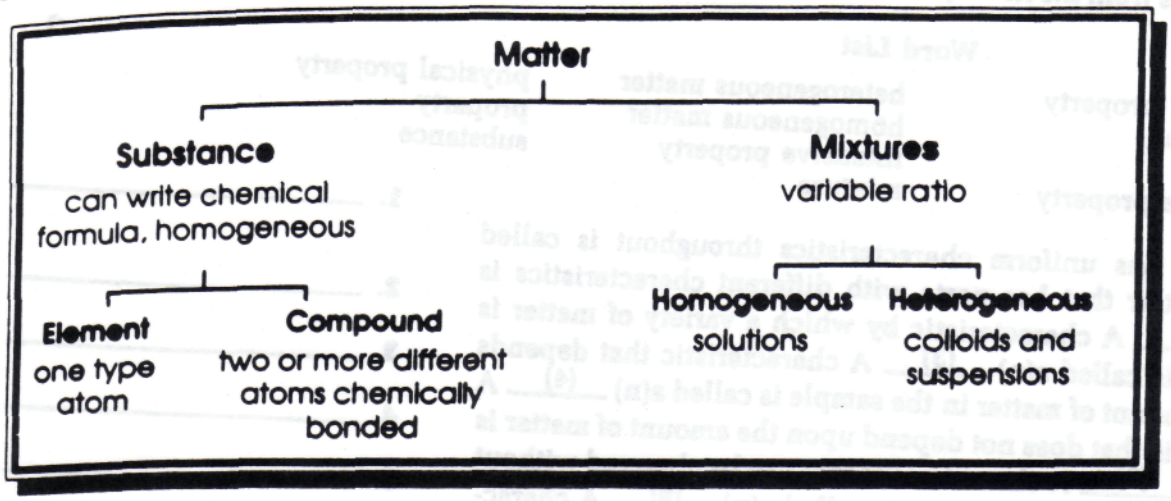
1. _____
2. _____
3. _____
4. _____
5. _____
6. _____
7. _____
8. _____
9. _____
10. _____
11. _____

*Classify each of the following as an **element**, **compound**, **heterogeneous mixture**, or **homogeneous mixture**.*

12. Water _____
13. Carbon _____
14. Air _____
15. Table salt _____
16. Sugar dissolved in water _____
17. Homogenized milk _____
18. Granite _____
19. Oxygen _____
20. Sand in water _____

MATTER — SUBSTANCES VS. MIXTURES

All matter can be classified as either a substance (element or compound) or a mixture (heterogeneous or homogeneous).



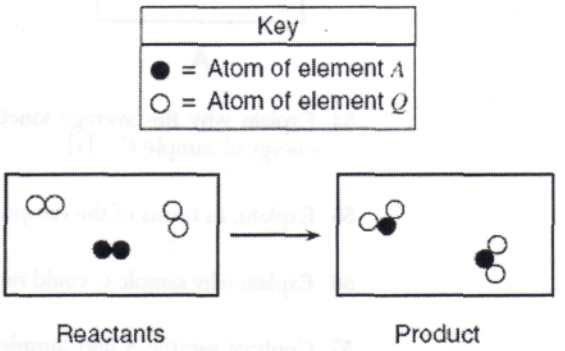
Classify each of the following as to whether it is a substance or a mixture. If it is a substance, write Element or Compound in the substance column. If it is a mixture, write Heterogeneous or Homogeneous in the mixture column.

Type of Matter	Substance	Mixture
1. chlorine		
2. water		
3. soil		
4. sugar water		
5. oxygen		
6. carbon dioxide		
7. rocky road ice cream		
8. alcohol		
9. pure air		
10. iron		

Practice with particle drawings

Base your answers to questions 55 through 57 on the information below.

The particle diagrams below represent the reaction between two nonmetals, A_2 and Q_2.

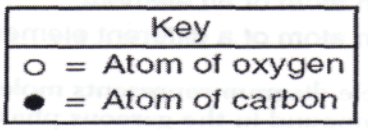


55 Using the symbols A and Q, write the chemical formula of the product. [1]

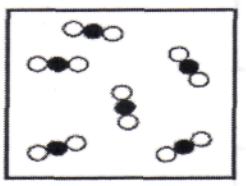
56 Identify the type of chemical bond between an atom of element A and an atom of element Q. [1]

57 Compare the total mass of the reactants to the total mass of the product. [1]

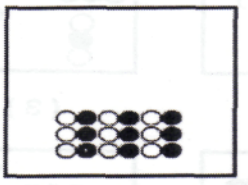
37 Given the key:



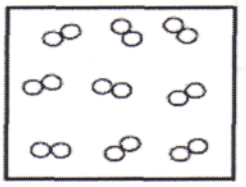
Which particle diagram represents a sample containing the compound CO(g)?



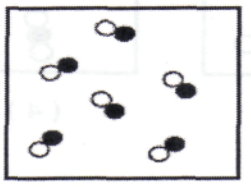
(1)



(3)

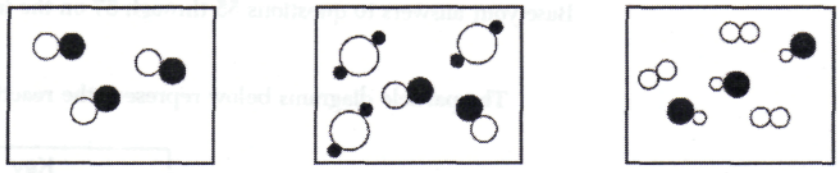


(2)



(4)

Base your answers to questions 54 through 57 on the particle diagrams below. Samples *A*, *B*, and *C* co[illegible] molecules at STP.



A B C

54 Explain why the average kinetic energy of sample *B* is equal to the average kinetic energy of sample *C*. [1]

55 Explain, in terms of the *composition*, why sample *A* represents a pure substance. [1]

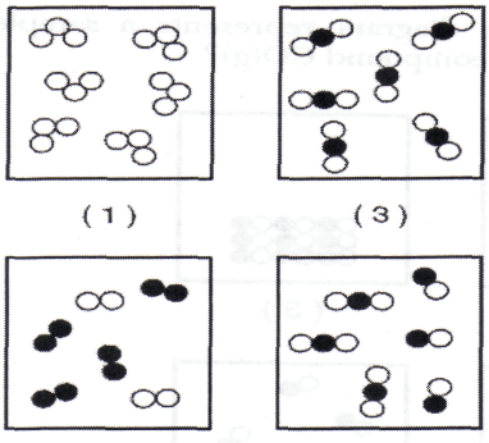
56 Explain why sample *C* could represent a mixture of fluorine and hydrogen chloride. [1]

57 Contrast sample *A* and sample *B*, in terms of *compounds and mixtures*. Include both sample *A* and sample *B* in your answer. [1]

43 Given the simple representations for atoms of two elements:

○ = an atom of an element
● = an atom of a different element

Which particle diagram represents molecules of only one compound in the gaseous phase?



(1) (3)

(2) (4)

51 In the boxes provided *in your answer booklet:*

a Draw *two* different compounds, one in each box, using the representations for atoms of element *X* and element *Z* given below. [1]

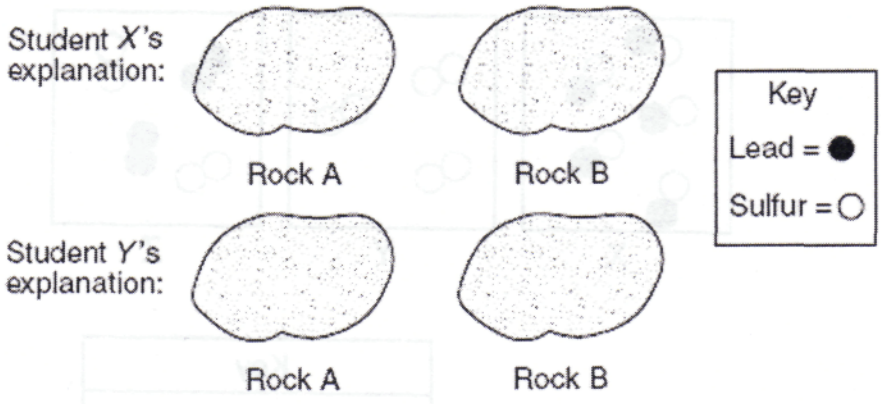
Atom of element *X* = ○
Atom of element *Z* = ●

b Draw a mixture of these two compounds. [1]

51 On a field trip, Student *X* and Student *Y* collected two rock samples. Analysis revealed that both rocks contained lead and sulfur. One rock contained a certain percentage of lead and sulfur by mass, and the other rock contained a different percentage of lead and sulfur by mass. Student *X* stated that the rocks contained two different mixtures of lead and sulfur. Student *Y* stated that the rocks contained two different compounds of lead and sulfur. Their teacher stated that both students could be correct.

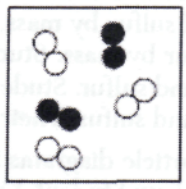
Draw particle diagrams in *each* of the rock diagrams provided *in your answer booklet* to show how Student *X*'s and Student *Y*'s explanations could both be correct. Use the symbols in the key provided *in your answer booklet* to sketch lead and sulfur atoms. [2]

Part B–2



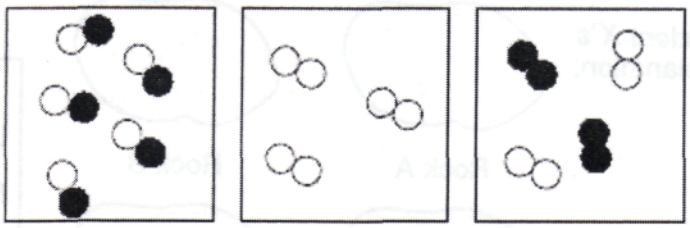
59 Given the reaction between two different elements in the gaseous state:

Box *A* below represents a mixture of the two reactants before the reaction occurs. The product of this reaction is a gas. In Box *B* provided *in your answer booklet*, draw the system after the reaction has gone to completion, based on the Law of Conservation of Matter. [2]



Box A
System Before Reaction

3 Given the diagrams *X*, *Y*, and *Z* below:



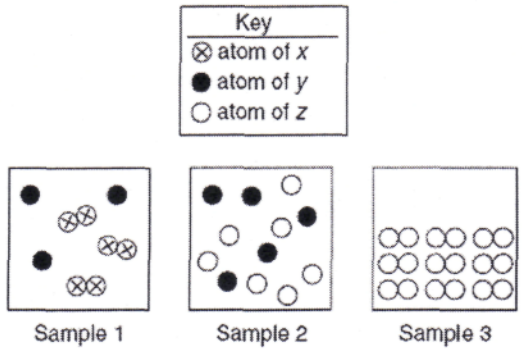
X Y Z

Key
Atom of element A = ○
Atom of element B = ●

Which diagram or diagrams represent a mixture of elements *A* and *B*?

(1) *X*, only
(2) *Z*, only
(3) *X* and *Y*
(4) *X* and *Z*

Base your answers to questions 65 through 67 on the particle diagrams below, which show atoms and/or molecules in three different samples of matter at STP.



65 Which sample represents a pure substance? [1]

66 When two atoms of y react with one atom of z, a compound forms. Using the number of atoms shown in sample 2, what is the maximum number of molecules of this compound that can be formed? [1]

67 Explain why ⊗⊗ does *not* represent a compound. [1]

_____

MATHEMATICS REVIEW FOR CHEMISTRY

1. METRIC SYSTEM AND MEASUREMENT:
   A. The metric system is used in chemistry is a DECIMAL system which has specific numerical relationship (multiples of ten) between units.
   B. The International Committees of Weights and Measures in France have standardized all the standard metric units. These units are called the International System of Units, Metric Units or SI units.
   C. The following is a list of some of the standard units that chemists use and what they measure: (refer to Table D on the reference table)

      Length—meter (m) Volume—liter (l) Mass— gram (:g)
      Time—sec Temperature—Kelvin

      Note: V= l x w x h
      V= cm x cm x cm
      V= cm^3
      cm^3= ml
      1000cm^3=1 liter

   D. Prefixes in the Metric System (refer to table C on the reference table)

Factor	Prefix	Symbol
10^{12}	Terra--	T
10^{9}	Giga--	G
10^{6}	Mega--	M
10^{3}	kilo--	k
10^{2}	hecto--	h
10^{1}	deka--	da,dk
10^{-1}	deci--	d
10^{-2}	centi--	c
10^{-3}	milli--	m
10^{-6}	micro--	μ
10^{-9}	nano	n
10^{-12}	pico	P

   E. Problem Solving: You may use any method you are comfortable with. I suggest you either shift the decimal point or do the factor label method. Try the following:

      1. 16 cm = _____ m
      2. 4.9 km = _____m
      3. 2.7 l = ____ml
      4. 7.8 g = ____mg

F. Conversion Factors:

Linear Measure—1 meter = 39.37 inches=1.0936 yards= 3.281 feet

Capacity—1 liter= o.9081 dry quart

Mass— 1 oz. = 28 g, 1 lb = 454 g, 1 kg = 2.2 lbs

Note: you have more of the smaller unit

Try:

1. 5 lbs = _____ g

2. 3 quarts = _____ L

3. 5 ft = _____ m

G. Scientific Notation: The numbers obtained in some measurements or calculations may be extremely small or large. In order to facilitate the writing of these numbers they are expressed as powers of ten.

For example:

1 000 000	1.0×10^{6}
175 000	1.75×10^{5}
0.000 056	5.6×10^{-5}

Try:

1. 61
2. .61
3. 610
4. .061
5. 6 100 000
6. .000 000 061

H. Mathematical Calculations with Exponential Numbers: Employ the following rules when using exponential numbers in mathematical operations. (Similar to algebra)

1. Addition and Subtraction: Numbers expressed in powers of 10 cannot be added or subtracted unless the powers of 10 are the same. If they are not the same, the numbers can be rewritten so that the powers of 10 are the same.

   Try:

   2.76×10^{5}
   $\underline{+\ 2.54 \times 10^{6}}$

   2.76×10^{-5}
   $\underline{-\ 2.54 \times 10^{-6}}$

2. Multiplication: When numbers with the power of 10 are multiplied, the exponents are added.
   Try: $(3.0 \times 10^{-2})(2.0 \times 10^{-4})$

3. Division: When divided the exponents are subtracted.
   Try: $(3.0 \times 10^{-2})/(1.5 \times 10^{-4})$

4. Square root: When taking the square root find the square root of the the number and divide the power of ten by 2.
   Try: Square root of 16×10^{8}

   Cube root of 8.0×10^{9}

I. Uncertainty in measurement:
   1. All measurements contain a certain degree of uncertainty because:
   a. the skill and carefulness of the person making the measurements
   b. the limitations of the measuring instrument
   2. Precision and accuracy: assuming that a skilled person uses the proper instrument the degree of certainty depends on two factors—precision and accuracy.
   a. Precision: indicates the reliability or reproducibility of a measurement
   b. Accuracy: indicates how close a measurement is to its accepted value.
   Try: Suppose you made several temperature measurements of boiling water using the same thermometer and your results are: 95.4, 95.2, 94.9, 95.1 and 94.8 are they precise and/ or accurate.

J. Significant figures: The accuracy of any measurement made in a chemistry Lab will depend on the precision of the instrument and the accuracy of the worker in using the instrument. Therefore when reporting a numerical measurement which has a limit of accuracy, it is customary to retain one doubtful figure. For example: 7.68 contain 3 significant figures—the first two ~~doubtful~~ certain, while the last one is doubtful.

   The following is a list of rules that you can use for determining the correct number of significant figures:
   1. Nonzero integers always count as significant figures.
   2. Zeros between figures are considered significant.
      For example: 705 have three significant figures
   3. Zeros, which appear in front of a number, are NOT considered significant.
      For example: 0.00876 contains three significant figures
   4. Zeros, which appear after a number are significant if followed by a decimal point or if the zeros are to the right of a decimal point.
      For example: 1600 has two significant figures
      1600. has four significant figures
      1600.00 has six significant figures

Try:

Number	Number of significant figure
66.6	
0.019	
1.01	
5500.7	
42.0	
0.700	
4.009	
0.008	
100.	
0.0550	

K. Numbers rounded off: In various calculations, the numerical answers can be rounded off to obtain the proper number of significant figures.

1. When the number dropped is 5 or greater the last significant figure is increased by 1.
   For example: 5.166 is changed to 5.17 so it contains 3 significant figures.
2. When the number dropped is less than 5 the last significant figure remains unchanged.
   For example: 5.164 is changed to 5.16 so that it contains 3 significant figures.
3. When adding and subtracting the answer should be rounded off so as to contain the least accurately known figure as the final one.
   For example:
   16.2 + 111.51 + 4.853 = 132.563 rounds to 132.6
   671.62-57.2 = 614.4 rounds to 614.4
4. When multiplying or dividing, the answer should be rounded so to contain only as many significant figures as are contained in the least accurate number.
   For example:
   4.123 x 5.12 = 21.10976 rounds off to 21.1
   6.0/2= 3.0 rounds to 3

Try :

REMEMBER... for addition and subtraction key is decimal place
For multiplication and division key is the # of s.f..

1. 2.032 + 1.1 + 3.68=
2. 16.0 + 33.=
3. 16.0 + 33.0 =
4. 52.5-36.25 =
5. 5.003 –2.15 =
6. 32.3 x 10. =
7. 32.3 x 10 =
8. 25 / 4.0 =

L. Percent error: % error = $\frac{\text{measured} - \text{accepted}}{\text{Accepted Value}} \times 100$

Table T

Accepted value: most probable value based on generally accepted reference
Experimental value: value based on laboratory reference
Try: If your laboratory measurement for the temperature of boiling water is 98 °C then your percent error is:

DIMENSIONAL ANALYSIS (FACTOR LABEL METHOD)

Name _____

Using this method, it is possible to solve many problems by using the relationship of one to another. For example, 12 inches = one foot. Since these two numbers represent the same value, the fractions 12 in/1 ft and 1 ft/12 in are both equal to one. When you multiply another number by the number one, you do not change its value. However, y may change its unit.

Example 1: Convert 2 miles to inches.

$$2 \text{ miles} \times \frac{5{,}280 \text{ ft}}{1 \text{ mile}} \times \frac{12 \text{ inches}}{1 \text{ ft}} = 126{,}720 \text{ in}$$

(Using significant figures, 2 mi = 100,000 in.)

Example 2: How many seconds are in 4 days?

$$4 \text{ days} \times \frac{24 \text{ hrs}}{1 \text{ day}} \times \frac{60 \text{ min}}{1 \text{ hr}} \times \frac{60 \text{ sec}}{1 \text{ min}} = 345{,}600 \text{ sec}$$

(Using significant figures, 4 days = 300,000 sec.)

Solve the following problems. Write the answers in significant figures.

1. 3 hrs = _____ sec
2. 0.035 mg = _____ cg
3. 5.5 kg = _____ lbs
4. 2.5 yds = _____ in
5. 1.3 yrs = _____ hr (1 yr = 365 days)
6. 3 moles = _____ molecules (1 mole = 6.02×10^{23} molecules)
7. 2.5×10^{24} molecules = _____ moles
8. 5 moles = _____ liters (1 mole = 22.4 liters)
9. 100. liters = _____ moles
10. 50. liters = _____ molecules
11. 5.0×10^{24} molecules = _____ liters
12. 7.5×10^{3} mL = _____ liters

METRICS AND MEASUREMENT

Name _____

In the chemistry classroom and lab, the metric system of measurement is used, so it is important to be able to convert from one unit to another.

mega	kilo	hecto	deca	**Basic Unit**	deci	centi	milli	micro
(M)	(k)	(h)	(da)	gram (g)	(d)	(c)	(m)	(μ)
1,000,000	1000	100	10	liter (L)	.1	.01	.001	.000001
10^6	10^3	10^2	10^1	meter (m)	10^{-1}	10^{-2}	10^{-3}	10^{-6}

Factor Label Method

1. Write the given number and unit.
2. Set up a conversion factor (fraction used to convert one unit to another).
   a. Place the given unit as denominator of conversion factor.
   b. Place desired unit as numerator.
   c. Place a "1" in front of the larger unit.
   d. Determine the number of smaller units needed to make "1" of the larger unit.
3. Cancel units. Solve the problem.

Example 1: 55 mm = _____ m

$$\frac{55\ \cancel{mm}}{} \cdot \frac{1\ m}{1000\ \cancel{mm}} = 0.055\ m$$

Example 2: 88 km = _____ m

$$\frac{88\ \cancel{km}}{} \cdot \frac{1000\ m}{1\ \cancel{km}} = 88{,}000\ m$$

Example 3: 7000 cm = _____ hm

$$\frac{7000\ \cancel{cm}}{} \cdot \frac{1\ \cancel{m}}{100\ \cancel{cm}} \cdot \frac{1\ hm}{100\ \cancel{m}} = 0.7\ hm$$

Example 4: 8 daL = _____ dL

$$\frac{8\ \cancel{daL}}{} \cdot \frac{10\ \cancel{L}}{1\ \cancel{daL}} \cdot \frac{10\ dL}{1\ \cancel{L}} = 800\ dL$$

The factor label method can be used to solve virtually any problem including changes in units. It is especially useful in making complex conversions dealing with concentrations and derived units.

Convert the following.

1. 35 mL = _____ dL
2. 950 g = _____ kg
3. 275 mm = _____ cm
4. 1,000 L = _____ kL
5. 1,000 mL = _____ L
6. 4,500 mg = _____ g
7. 25 cm = _____ mm
8. 0.005 kg = _____ dag
9. 0.075 m = _____ cm
10. 15 g = _____ mg

SCIENTIFIC NOTATION

Name _____

Scientists very often deal with very small and very large numbers, which can lead to a lot of confusion when counting zeros! We have learned to express these numbers as pow of 10.

Scientific notation takes the form of $M \times 10^n$ where $1 \leq M < 10$ and "n" represents the number of decimal places to be moved. Positive n indicates the standard form is larger than zero whereas negative n would indicate a number smaller than zero.

Example 1: Convert 1,500,000 to scientific notation. We move the decimal point so that there is only one digit to its left, a total of 6 places.

$1{,}500{,}000 = 1.5 \times 10^6$

Example 2: Convert 0.000025 to scientific notation. For this, we move the decimal point 5 places to the right.

$0.000025 = 2.5 \times 10^{-5}$

(Note that when a number starts out less than one, the exponent is always negative.)

Convert the following to scientific notation.

1. 0.005 = _____
2. 5,050 = _____
3. 0.0008 = _____
4. 1,000 = _____
5. 1,000,000 = _____
6. 0.25 = _____
7. 0.025 = _____
8. 0.0025 = _____
9. 500 = _____
10. 5,000 = _____

Convert the following to standard notation.

1. 1.5×10^3 = _____
2. 1.5×10^{-3} = _____
3. 3.75×10^{-2} = _____
4. 3.75×10^2 = _____
5. 2.2×10^5 = _____
6. 3.35×10^{-1} = _____
7. 1.2×10^{-4} = _____
8. 1×10^4 = _____
9. 1×10^{-1} = _____
10. 4×10^0 = _____

SIGNIFICANT FIGURES

Name _____

A measurement can only be as accurate and precise as the instrument that produced it. A scientist must be able to express the accuracy of a number, not just its numerical value. We can determine the accuracy of a number by the number of significant figures it contains.

1) All digits 1-9 inclusive are significant.
   Example: 129 has 3 significant figures.
2) Zeros between significant digits are always significant.
   Example: 5,007 has 4 significant figures.
3) Trailing zeros in a number are significant <u>only</u> if the number contains a decimal point.
   Example: 100.0 has 4 significant figures.
   100 has 1 significant figure.
4) Zeros in the beginning of a number whose only function is to place the decimal point are not significant.
   Example: 0.0025 has 2 significant figures.
5) Zeros following a decimal significant figure are significant.
   Example: 0.000<u>470</u> has 3 significant figures.
   0.<u>47000</u> has 5 significant figures.

Determine the number of significant figures in the following numbers.

1. 0.02 _____
2. 0.020 _____
3. 501 _____
4. 501.0 _____
5. 5,000 _____
6. 5,000. _____
7. 6,051.00 _____
8. 0.0005 _____
9. 0.1020 _____
10. 10,001 _____

Determine the location of the last significant place value by placing a bar over the digit. (Example: $1.70\bar{0}$)

1. 8040 _____
2. 0.0300 _____
3. 699.5 _____
4. 2.000×10^2 _____
5. 0.90100 _____
6. 90,100 _____
7. 4.7×10^{-8} _____
8. 10,800,000. _____
9. 3.01×10^{21} _____
10. 0.000410 _____

CALCULATIONS USING SIGNIFICANT FIGURES

Name _____

When multiplying and dividing, limit and round to the least number of significant figures any of the factors.

Example 1: 23.0 cm x 432 cm x 19 cm = 188,784 cm^3
The answer is expressed as 190,000 cm^3 since 19 cm has only two significant figures.

When adding and subtracting, limit and round your answer to the least number of decir places in any of the numbers that make up your answer.

Example 2: 123.25 mL + 46.0 mL + 86.257 mL = 255.507 mL
The answer is expressed as 255.5 mL since 46.0 mL has only one decimal place.

Perform the following operations expressing the answer in the correct number of significc figures.

1. 1.35 m x 2.467 m = _____
2. 1,035 m^2 ÷ 42 m = _____
3. 12.01 mL + 35.2 mL + 6 mL = _____
4. 55.46 g − 28.9 g = _____
5. .021 cm x 3.2 cm x 100.1 cm = _____
6. 0.15 cm + 1.15 cm + 2.051 cm = _____
7. 150 L^3 ÷ 4 L = _____
8. 505 kg − 450.25 kg = _____
9. 1.252 mm x 0.115 mm x 0.012 mm = _____
10. 1.278×10^3 m^2 ÷ 1.4267×10^2 m = _____

PERCENTAGE ERROR

Name _____

Percentage error is a way for scientists to express how far off a laboratory value is from the commonly accepted value.

The formula is:

$$\% \text{ error} = \frac{\text{measured} - \text{accepted}}{\text{Accepted Value}} \times 100$$

Determine the percentage error in the following problems.

1. Experimental Value = 1.24 g
   Accepted Value = 1.30 g

   Answer: _____

2. Experimental Value = 1.24×10^{-2} g
   Accepted Value = 9.98×10^{-3} g

   Answer: _____

3. Experimental Value = 252 mL
   Accepted Value = 225 mL

   Answer: _____

4. Experimental Value = 22.2 L
   Accepted Value = 22.4 L

   Answer: _____

5. Experimental Value = 125.2 mg
   Accepted Value = 124.8 mg

   Answer: _____