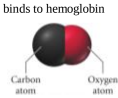
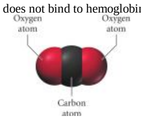
Chapter 1 Matter, Measurement, and Problem Solving

Structure Determines Properties

- the properties of matter are determined by the atoms and molecules that compose it

carbon monoxide	carbon dioxide
1. composed of one carbon atom and one oxygen atom	1. composed of one carbon atom and two oxygen atoms
2. colorless, odorless gas	2. colorless, odorless gas
3. burns with a blue flame	3. incombustible
4. binds to hemoglobin	4. does not bind to hemoglobin





Atoms and Molecules

- **atoms**
  - ✓ are submicroscopic particles
  - ✓ are the fundamental building blocks of all matter
- **molecules**
  - ✓ two or more atoms attached together
    - ➢ attachments are called bonds
    - ➢ attachments come in different strengths
  - ✓ molecules come in different shapes and patterns
- **Chemistry** is the science that seeks to understand the behavior of matter by studying the behavior of atoms and molecules

ELEMENTS to MEMORIZE

Aluminum	**Al**	**Manganese**	**Mn**
Antimony	**Sb**	**Mercury**	**Hg**
Argon	**Ar**	**Neon**	**Ne**
Arsenic	**As**	**Nickel**	**Ni**
Barium	**Ba**	**Nitrogen**	**N**
Beryllium	**Be**	**Oxygen**	**O**
Boron	**B**	**Palladium**	**Pd**
Bromine	**Br**	**Phosphorus**	**P**
Calcium	**Ca**	**Platinum**	**Pt**
Carbon	**C**	**Plutonium**	**Pu**
Cesium	**Cs**	**Potassium**	**K**
Chlorine	**Cl**	**Radium**	**Ra**
Chromium	**Cr**	**Radon**	**Rn**
Cobalt	**Co**	**Rubidium**	**Rb**
Copper	**Cu**	**Selenium**	**Se**
Fluorine	**F**	**Silicon**	**Si**
Gallium	**Ga**	**Silver**	**Ag**
Germanium	**Ge**	**Sodium**	**Na**
Gold	**Au**	**Strontium**	**Sr**
Helium	**He**	**Sulfur**	**S**
Hydrogen	**H**	**Tin**	**Sn**
Iodine	**I**	**Titanium**	**Ti**
Iron	**Fe**	**Tungsten**	**W**
Krypton	**Kr**	**Uranium**	**U**
Lead	**Pb**	**Xenon**	**Xe**
Lithium	**Li**	**Zinc**	**Zn**
Magnesium	**Mg**	**Zirconium**	**Zr**

The Scientific Approach to Knowledge

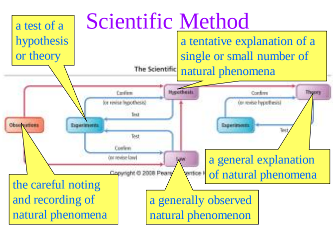
- philosophers try to understand the universe by reasoning and thinking about "ideal" behavior
- scientists try to understand the universe through empirical knowledge gained through observation and experiment

SCIENTIFIC METHOD

1. **FACT:** An observable event; indisputable evidence which does not explain but simply is.
2. **HYPOTHESIS:** A guess to try to explain an observation.
3. **EXPERIMENT:** A systematic exploration of an observation or concept.
4. **THEORY:** An explanation of the facts; it can be proven by experiment and it confirms an hypothesis.
5. **LAW:** A theory which has undergone rigorous experimentation and no contradiction can be found.

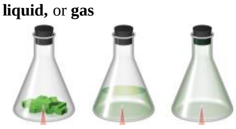
Note: **MODEL:** A visual or mathematical device or method used to demonstrate a theory or concept.

Scientific Method



Classification of Matter

- **matter** is anything that has mass and occupies space
- we can classify matter based on whether it's **solid, liquid,** or **gas**



Classifying Matter by Physical State

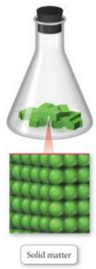
- matter can be classified as solid, liquid, or gas based on the characteristics it exhibits

State	*Shape*	*Volume*	*Compress*	*Flow*
Solid	Fixed	Fixed	No	No
Liquid	Indef.	Fixed	No	Yes
Gas	Indef.	Indef.	Yes	Yes

- Fixed = keeps shape when placed in a container
- Indefinite = takes the shape of the container

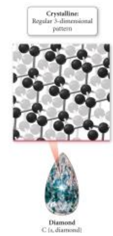
Solids

- the particles in a solid are packed close together and are fixed in position
  - ✓ though they may vibrate
- the close packing of the particles results in solids being incompressible
- the inability of the particles to move around results in solids retaining their shape and volume when placed in a new container, and prevents the particles from flowing



Crystalline Solids

- some solids have their particles arranged in an orderly geometric pattern – we call these **crystalline solids**
  - ✓ salt and diamonds

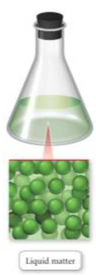


Amorphous Solids

- some solids have their particles randomly distributed without any long-range pattern – we call these **amorphous solids**
  - ✓ plastic
  - ✓ glass
  - ✓ charcoal

Liquids

- the particles in a liquid are closely packed, but they have some ability to move around
- the close packing results in liquids being incompressible
- but the ability of the particles to move allows liquids to take the shape of their container and to flow – however, they don't have enough freedom to escape and expand to fill the container



Gases

- in the gas state, the particles have complete freedom from each other
- the particles are constantly flying around, bumping into each other and the container
- in the gas state, there is a lot of empty space between the particles
  - ✓ on average

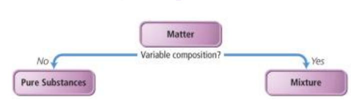
Gases

- because there is a lot of empty space, the particles can be squeezed closer together – therefore gases are compressible
- because the particles are not held in close contact and are moving freely, gases expand to fill and take the shape of their container, and will flow

Classification of Matter by Composition

- matter whose composition does not change from one sample to another is called a **pure substance**
  - ✓ made of a single type of atom or molecule
  - ✓ because composition is always the same, all samples have the same characteristics
- matter whose composition may vary from one sample to another is called a **mixture**
  - ✓ two or more types of atoms or molecules combined in variable proportions
  - ✓ because composition varies, samples have the different characteristics

Classification of Matter by Composition



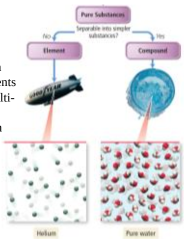
Pure Substances:
1) made of one type of particle
2) all samples show the same intensive properties

Mixture:
1) made of multiple types of particles
2) samples may show different intensive properties

Classification of Pure Substances

- substances that cannot be broken down into simpler substances by chemical reactions are called **elements**
  - ✓ basic building blocks of matter
  - ✓ composed of single type of atom
    - ➢ though those atoms may or may not be combined into molecules
- substances that can be decomposed are called **compounds**
  - ✓ chemical combinations of elements
  - ✓ composed of molecules that contain two or more different kinds of atoms
  - ✓ all molecules of a compound are identical, so all samples of a compound behave the same way
- most natural pure substances are compounds

Classification of Pure Substances



Element:
1) made of one type of atom (some elements found as multi-atom molecules in nature)
2) combine together to make compounds

Compound:
1) made of one type of molecule, or array of ions
2) molecules contain 2 or more different kinds of atoms

Classification of Mixtures

- **homogeneous** = mixture that has uniform composition throughout
  - ✓every piece of a sample has identical characteristics, though another sample with the same components may have different characteristics
  - ✓atoms or molecules mixed uniformly
- **heterogeneous** = mixture that does not have uniform composition throughout
  - ✓contains regions within the sample with different characteristics
  - ✓atoms or molecules not mixed uniformly

Classification of Mixtures

Heterogeneous:
1) made of multiple substances, whose presence can be seen
2) portions of a sample have different composition and properties

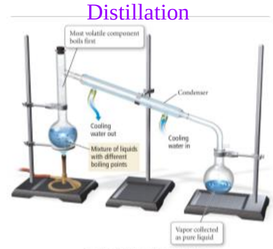
Homogeneous:
1) made of multiple substances, but appears to be one substance
2) all portions of a sample have the same composition and properties

Separation of Mixtures

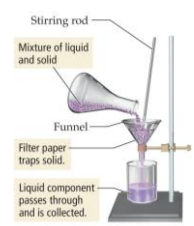
- separate mixtures based on different physical properties of the components
  - ✓Physical change

Different Physical Property	Technique
Boiling Point	Distillation
State of Matter (solid/liquid/gas)	Filtration
Adherence to a Surface	Chromatography
Volatility	Evaporation
Density	Centrifugation & Decanting

Distillation



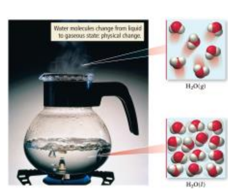
Filtration



Changes in Matter

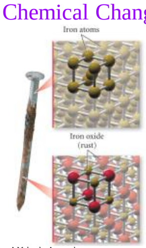
- changes that alter the state or appearance of the matter without altering the composition are called **physical changes**
- changes that alter the composition of the matter are called **chemical changes**
  - ✓during the chemical change, the atoms that are present rearrange into new molecules, but all of the original atoms are still present

Physical Changes in Matter



The boiling of water is a physical change. The water molecules are separated from each other, but their structure and composition do not change.

Chemical Changes in Matter



The rusting of iron is a chemical change. The iron atoms in the nail combine with oxygen atoms from O_2 in the air to make a new substance, rust, with a different composition.

Properties of Matter

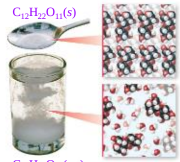
- **physical properties** are the characteristics of matter that can be changed without changing its composition
  - ✓characteristics that are directly observable
- **chemical properties** are the characteristics that determine how the composition of matter changes as a result of contact with other matter or the influence of energy
  - ✓characteristics that describe the behavior of matter

Common Physical Changes

- processes that cause changes in the matter that do not change its composition
- state changes
  - ✓boiling / condensing
  - ✓melting / freezing
  - ✓subliming
- dissolving

Dissolving of Sugar / Subliming of Dry Ice

$C_{12}H_{22}O_{11}(s)$



$C_{12}H_{22}O_{11}(aq)$

Common Chemical Changes

- processes that cause changes in the matter that change its composition
- rusting
- processes that release lots of energy
- burning

$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(l)$

Energy Changes in Matter

- changes in matter, both physical and chemical, result in the matter either gaining or releasing energy
- **energy** is the capacity to do work
- work is the action of a force applied across a distance
  - ✓a force is a push or a pull on an object
  - ✓electrostatic force is the push or pull on objects that have an electrical charge

Energy of Matter

- all matter possesses energy
- energy is classified as either kinetic or potential
- energy can be converted from one form to another
- when matter undergoes a chemical or physical change, the amount of energy in the matter changes as well

Energy of Matter - Kinetic

- **kinetic energy** is energy of motion
  - ✓motion of the atoms, molecules, and subatomic particles
  - ✓thermal (heat) energy is a form of kinetic energy because it is caused by molecular motion

Energy of Matter - Potential

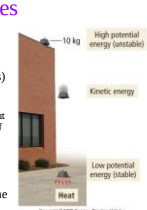
- **potential energy** is energy that is stored in the matter
  - ✓due to the composition of the matter and its position in the universe
  - ✓chemical potential energy arises from electrostatic forces between atoms, molecules, and subatomic particles

Conversion of Energy

- you can interconvert kinetic energy and potential energy
- whatever process you do that converts energy from one type or form to another, the total amount of energy remains the same
  - ✓**Law of Conservation of Energy**

Spontaneous Processes

- materials that possess high potential energy are less stable
- processes in nature tend to occur on their own when the result is material(s) with lower total potential energy
  - ✓ processes that result in materials with higher total potential energy can occur, but generally will not happen without input of energy from an outside source
- when a process results in materials with less potential energy at the end than there was at the beginning, the difference in energy is released into the environment



WORKSHOP - Matter

1. From the thermal decomposition of a pure solid, we obtained a solid and a gas, each of which is a pure substance. From this information, we can conclude with certainty that
   a. the original solid is not an element
   b. both products are elements
   c. at least one of the products is an element
   d. the solid is a compound and the gas is an element

2. A solution can be distinguished from a compound by its
   a) liquid state
   b) heterogeneous nature
   c) lack of color
   d) variable composition
   e) all of the above

3. List the name and symbols for the first 80 elements.

4. Distinguish the chemical properties from the physical properties of copper? Give examples.

5. A clear blue liquid in an open beaker was left in the hood. After 1 week, the beaker contained only blue crystals. The original liquid can be classified as a(n) _______________.

6. Describe the best technique for separating a solid from its liquid. What techniques would you use to separate with liquids?

Standard Units of Measure

MEASUREMENTS
Scientific Notation

Many measurements in science involve either very large numbers or very small numbers (#). Scientific notation is one method for communicating these types of numbers with minimal writing.

GENERIC FORMAT: # . # #… x $10^{\#}$

A negative exponent represents a number less than 1 and a positive exponent represents a number greater than 1.

6.75×10^{-3} is the same as 0.00675
6.75×10^{3} is the same as 6750

MEASUREMENTS
Scientific Notation Practice

Give the following in scientific notation (or write it out) with the appropriate significant figures.

1. **528900300000 =** 5.289003×10^{11}
2. **0.000000000003400 =** 3.400×10^{-12}
3. **0.23 =** 2.3×10^{-1}
4. **5.678×10^{-7} =** 0.0000005678
5. **9.8×10^{4} =** 98000

MEASUREMENTS
Significant Figures

I. All nonzero numbers are significant figures.
II. Zero's follow the rules below.
1. **Zero's between numbers are significant.**
   30.09 has 4 SF
2. **Zero's that precede are NOT significant.**
   0.000034 has 2 SF
3. **Zero's at the end of decimals are significant.**
   0.00900 has 3 SF
4. **Zero's at the end without decimals are either.**
   4050 has either 4 SF or 3 SF

MEASUREMENTS
Significant Figures & Calculations

Significant figures are based on the tools used to make the measurement. An imprecise tool will negate the precision of the other tools used. The following rules are used when measurements are used in calculations.

Adding/subtracting:

The result should be rounded to the same number of decimal places as the measurement with the least decimal places.

Multiplying/dividing:

The result should contain the same number of significant figures as the measurement with the least significant figures.

Multiplication and Division with Significant Figures

- when multiplying or dividing measurements with significant figures, the result has the same number of significant figures as the measurement with the fewest number of significant figures

5.02 × 89,665 × 0.10 = 45.0118 = 45
3 sig. figs. 5 sig. figs. 2 sig. figs. 2 sig. figs.

5.892 ÷ 6.10 = 0.96590 = 0.966
4 sig. figs. 3 sig. figs. 3 sig. figs.

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Addition and Subtraction with Significant Figures

- when adding or subtracting measurements with significant figures, the result has the same number of decimal places as the measurement with the fewest number of decimal places

5.74 + 0.823 + 2.651 = 9.214 = 9.21
2 dec. pl. 3 dec. pl. 3 dec. pl. 2 dec. pl.

4.8 - 3.965 = 0.835 = 0.8
1 dec. pl 3 dec. pl. 1 dec. pl.

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MEASUREMENTS
Significant Figures & Calculations

2 Answers are Incorrect!!!

Adding & Subtracting

345.678	0.07283	1587
+ 12.67	- 0.0162789	- 120
358.348	0.0565511	1467
358.35	0.05655	If 3 SF: 1470 or 1.47×10^{3}

Multiplication & Division

(12.034)(3.98) =	47.89532	47.9 is correct
98.657 ÷ 43 =	2.294348837	2.3 is correct
$\frac{(13.59)(6.3)}{12}$ =	7.13475	7.1 is correct

PRACTICE PROBLEMS

Show your work for the following questions on the back. Always give the correct significant figures.

1. Express each of the following numbers in scientific notation & 3 significant figures.

A) 6545490087 **6.55×10^{9}**　　C) 0.0002368 **2.37×10^{-4}**

B) 0.000001243 **1.24×10^{-6}**　　D) 94560 **9.46×10^{4}**

2. 0.00496 - 0.00298 = **1.98×10^{-3}**

3. (3.36-5.6) / (82.98 + 2.4) = **-2.6×10^{-2}**

4. $4.45 \times 10^{-23} / 8.345 \times 10^{-53}$ = **5.33×10^{29}**

5. $[(26.7 \times 10^{-8})(47 \times 10^{13})]^4 / (8.54 \times 10^{17})^{1/2}$ = **2.7×10^{23}**

DIMENSIONAL ANALYSIS

Unit Conversions

Common SI Prefixes:

Factor	Prefix	Abbreviation
10^{6}	Mega	M
10^{3}	Kilo	k
10^{2}	Hecto	h
10^{1}	Deka	da
10^{-1}	Deci	d
10^{-2}	Centi	c
10^{-3}	Milli	m
10^{-6}	Micro	µ
10^{-9}	Nano	n
10^{-12}	Pico	p

The Standard Units

- Scientists have agreed on a set of international standard units for comparing all our measurements called the SI units
  - ✓*Système International* = International System

Quantity	Unit	Symbol
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K

TEMPERATURE CONVERSIONS

1. Fahrenheit – at standard atmospheric pressure, the melting point of ice is 32 °F, the boiling point of water is 212 °F, and the interval between is divided into 180 equal parts.

2. Celsius – at standard atmospheric pressure, the melting point of ice is 0 °C, the boiling point of water is 100 °C, and the interval between is divided into 100 equal parts.

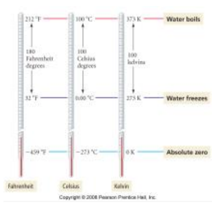
3. Kelvin – assigns a value of zero to the lowest conceivable temperature; there are NO negative numbers.

$$T(K) = T(°C) + 273.15$$

$$T(°F) = 1.8T(°C) + 32$$

Temperature Scales

- Fahrenheit Scale, °F
  - ✓ used in the U.S.
- Celsius Scale, °C
  - ✓ used in all other countries
- Kelvin Scale, K
  - ✓ absolute scale
    - ➢ no negative numbers
  - ✓ directly proportional to average amount of kinetic energy
  - ✓ 0 K = absolute zero



Dimensional Analysis

Dimensional Analysis (also call unit analysis) is one method for solving math problems that involve measurements. The basic concept is to use the units associated with the measurement when determining the next step necessary to solve the problem. Always start with the given measurement then immediately follow the measurement with a set of parentheses.

Keep in mind, try to ask yourself the following questions in order to help yourself determine what to do next.

1. **Do I want that unit?**

   If not, get rid of it by dividing by it if the unit is in the numerator, (if the unit is in the denominator, then multiply).

2. **What do I want?**

Place the unit of interest in the opposite position in the parentheses.

Numerator
Denominator

MEASUREMENTS LECTURE - METRIC

1. How many meters are equal to 16.80 km?

2. How many cubic centimeters are there in 1 cubic meter?

3. How many nm are there in 200 dm? Express your answer in scientific notation.

4. How many mg are there in 0.5 kg?

MEASUREMENTS PRACTICE - METRIC

1. The mass of a young student is found to be 87 kg. How many grams does this mass correspond to?

87 kg (1000g / 1 kg) = 87000 g or 8.7×10^{4} g

2. How many liters is equivalent to 15.0 cubic meters?

15.0 m^3 (100 cm / 1 m)^3 (1 mL/1 cm^3) (1 L/1000 mL) = 1.50×10^{4} L

MEASUREMENTS

Since two different measuring systems exist, a scientist must be able to convert from one system to the other.

CONVERSIONS

Length:	**1 in = 2.54 cm**	**1 mi = 1.61 km**
Mass:	**1 lb = 454 g**	**1 kg = 2.2 lb**
Volume:	**1 qt = 946 mL**	**1 L = 1.057 qt**
	4 qt = 1 gal	**1 mL = 1 cm^3**
Temperature:	**°F = (1.8 °C) + 32**	
	°C = $\frac{(°F - 32)}{1.8}$	**K = °C + 273.15**

MEASUREMENTS LECTURE - CONVERSIONS

1. The mass of a young student is found to be 87 kg. How many pounds does this mass correspond to?

2. An American visited Austria during the summer summer, and the speedometer in the taxi read 90 km/hr. How fast was the American driving in miles per hour? ***(Note: 1 mile = 1.6093 km)***

3. In most countries, meat is sold in the market by the kilogram. Suppose the price of a certain cut of beef is 1400 pesos/kg, and the exchange rate is 124 pesos to the U.S. dollar. What is the cost of the meat in dollars per pound (lb)? ***(Note: 1 kg = 2.20 lb)***

PRACTICE PROBLEMS

1. **Convert 15.0 J to kcal** **0.00359 kcal**
2. **Convert 15.0 mg to pounds** **3.30×10^{-5} lb**
3. **Convert 15.0 ft^3 to cL** **4.25×10^4 cL**
4. **How many liters of gasoline will be used to drive 725 miles in a car that averages 27.8 miles per gallon?** **98.7 L**
5. **Diamonds crystallize directly from rock melts rich in magnesium and saturated carbon dioxide gas that has been subjected to high pressures and temperatures exceeding 1677 K. Calculate this temperature in Fahrenheit.** **2559 °F**
6. **D.J. promised to bake 25 dozen cookies and deliver them to a bake sale. If each cookie weighs 3.5 ounces, how many kilograms will 25 dozen cookies weigh?** **30. kg**

WORKSHOP 1 Measurements

1. The melting point (freezing point) of mercury is -35°C. What is the melting point temperature, in degrees Fahrenheit? Normal body temperature is 98.6°F. What is this in Celsius?

2. Which is the shortest distance?
   - a) 100,000 millimeters
   - b) 1×10^{-3} kilometers
   - c) 100,000,000 micrometers
   - d) 1,000 centimeters
   - e) 1×10^{-4} mega meters
   - f) 100×10^{10} nm

3a. When a weight of 28.53 grams is added to a weight of 1.792 grams, the result should be expressed as: _________________ (sig fig/sci not/units)

3b. When a weight of 28.5345 grams is divided by 1.792 mL, the result should be expressed as: ____________________ (sig fig/sci not/units)

4. How many liters of gasoline will be used to drive 103 km in a car that averages 25.8 miles per gallon?

5. Calculate the number of weeks in 672 hours.

6. Fluorite, a mineral of calcium, is a compound of the metal with fluoride. Analysis shows that a 0.01215 lb sample of fluorite contains 2840 mg of calcium. Calculate the mass percent of calcium and fluoride in fluorite.

Density

INTRODUCTION TO DENSITY

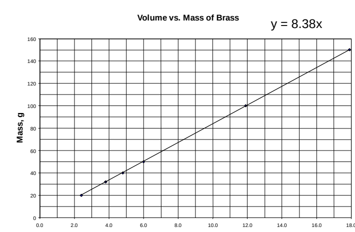
- Density is the measurement of the mass of an object per unit volume of that object.

  d = m / V

- Density is usually measured in g/mL or g/cm^3 for solids or liquids.
- Volume may be measured in the lab using a graduated cylinder or calculated using:

  Volume = length x width x height if a box or
  $V = \pi r^2 h$ if a cylinder.

- Remember 1 mL = 1 cm^3



DENSITY DETERMINATION

1. Mercury is the only metal that is a liquid at 25 °C. Given that 1.667 mL of mercury has a mass of 22.60 g at 25 °C, calculate its density.

2. Iridium is a metal with the greatest density, 22.65 g/cm³. What is the volume of 192.2 g of Iridium?

3. What volume of acetone has the same mass as 10.0 mL of mercury? Take the densities of acetone and mercury to be 0.792 g/cm³ and 13.56 g/cm³, respectively.

4. Hematite (iron ore) weighing 70.7 g was placed in a flask whose volume was 53.2 mL. The flask was then carefully filled with water and weighed. Hematite and water combined weighed 109.3 g. The density of water is 0.997 g/cm³. What is the density of hematite?

PRACTICE PROBLEMS

A study of gemstones and dimensional analysis:

The basic unit for gemstones is the carat. One carat is equal to 200 milligrams. A well-cut diamond of one carat measures 0.25 inches exactly in diameter. Right click for answers

1.13×10^5 mg **1. The Star of India sapphire (Al_2O_3, corundum) weighs 563 carats. What is the weight of the gemstone in milligrams?**

621.2 g **2. The world's largest uncut diamond (C, an allotrope of carbon) was the Cullinan Diamond. It was discovered 1/25/1905 in Transvaal, South Africa. It weighed 3,106 carats. Calculate this weight in grams.**

0.3948 kg **3. The Cullinan Diamond was cut into nine major stones and 96 smaller brilliants. The total weight of the cut stones was 1063 carats, only 35% of the original weight! What weight (in kilograms) of the Cullinan Diamond was not turned into gemstones?**

0.54696 m **4. Emerald is a variety of green beryl ($Be_3Al_2Si_6O_{18}$) that is colored by a trace of chromium, which replaces aluminum in the beryl structure. The largest cut emerald was found in Carnaiba, Brazil Aug. 1974. It weighs 86,136 carats. Assuming the diamond carat to size relationship stands for emeralds, calculate the approximate diameter of this stone in meters.**

0.59 L **5. The largest cut diamond, the Star of Africa, is a pear-shaped diamond weighing 530.2 carats. It is 2.12 in long, 4.4 cm wide, and 250 mm thick at its deepest point. What is the minimum volume (in liters) of a box that could be used to hide this diamond.**

WORKSHOP 1Density

1. For each of the following cases, state whether the density of the object increases (I), decreases (D), or remains the same (R).
   - **a) ________A sample of chlorine gas is compressed.**
   - **b) ________A lead weight is carried from sea level to the top of a high mountain.**
   - **c) ________A sample of liquid water is frozen.**

2. Calculate the density of a box that is 750 mm long by 55 cm wide by 9 in high and weighs 24.33g. Will it sink or float in water? Justify your answer mathematically.

3. Water, with a density of 1.00 g/ml and ethyl alcohol, with a density of 0.789g/ml, are combined to form a mixture which is 50% by weight of each substance. If 56 grams of water are used, how many milliliters of the alcohol must be used?

4. A sample of metallic element X, weighing 3.177 g, combines completely with 0.6015 L of O_2 gas (at ambient temperature and pressure) to form the metal oxide with the formula XO. If the density of O_2 gas is 1.330 g/L, identify X?

Precision and Accuracy

Uncertainty in Measured Numbers

- uncertainty comes from limitations of the instruments used for comparison, the experimental design, the experimenter, and nature's random behavior
- to understand how reliable a measurement is we need to understand the limitations of the measurement
- **accuracy** is an indication of how close a measurement comes to the **actual** value of the quantity
- **precision** is an indication of how reproducible a measurement is

Precision

- imprecision in measurements is caused by **random errors**
  - ✓errors that result from random fluctuations
  - ✓no specific cause, therefore cannot be corrected
- we determine the precision of a set of measurements by evaluating how far they are from the actual value and each other
- even though every measurement has some random error, with enough measurements these errors should average out

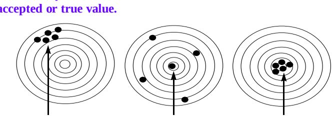
Accuracy

- inaccuracy in measurement caused by **systematic errors**
  - ✓errors caused by limitations in the instruments or techniques or experimental design
  - ✓can be reduced by using more accurate instruments, or better technique or experimental design
- we determine the accuracy of a measurement by evaluating how far it is from the actual value
- systematic errors do not average out with repeated measurements because they consistently cause the measurement to be either too high or too low

PRECISION AND ACCURACY

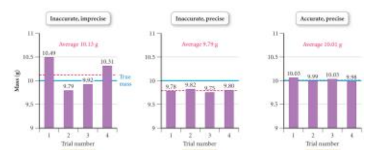
1. Precision – refers to the degree of reproducibility of a measured quantity.

2. Accuracy – refers to how close a measured value is to the accepted or true value.



Precise (not accurate) Accurate (not precise) Both Precise/Accurate

Accuracy vs. Precision



STANDARD DEVIATION

The **standard deviation** of a series of measurements which includes at least 6 independent trials may be defined as follows. If we let $\mathbf{x}_m$ be a measured value, **N** be the number of measurements, $\langle\mathbf{x}\rangle$ be the average or **mean** of all the measurements, then **d** is the **deviation** of a value from the average:

$$\mathbf{d} = \mathbf{x}_m - \langle\mathbf{x}\rangle$$

and the standard deviation, **s**, is defined by:

$$s = \sqrt{\frac{\sum d^2}{(N-1)}}$$

where $\Sigma \mathbf{d}^2$ means "sum of all the values of $\mathbf{d}^2$."

The **value of the measurement** should include some indication of the precision of the measurement. The standard deviation is used for this purpose if a large number of measurements of the same quantity is subject to random errors only. We can understand the meaning of **s** if we plot on the y-axis the number of times a given value of $\mathbf{x}_m$ is obtained, against the values, $\mathbf{x}_m$, on the x-axis. The "normal distribution curve" is bell-shaped, with the most frequent value being the average value, $\langle\mathbf{x}\rangle$.

STANDARD DEVIATION

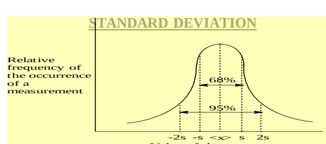


Figure 3: Distribution of Values of a Measurement

Most of the measurements give values near $\langle\mathbf{x}\rangle$. In fact, 68% of the measurements fall within the standard deviation **s** of $\langle\mathbf{x}\rangle$ (see graph). 95% of the measured values are found within 2**s** of $\langle\mathbf{x}\rangle$. We call the value of 2**s** the uncertainty of the measurement, **u**. Then, if we report our value of the measurement as $\langle\mathbf{x}\rangle \pm \mathbf{u}$, we are saying that $\langle\mathbf{x}\rangle$ is the most probable value and 95% of the measured values fall within this range. The next example shows how the standard deviation can be used to evaluate the data.

STANDARD DEVIATION

Example 1. Weight of a test tube on 10 different balances

trial	weight	$d = X_m - \langle X\rangle$	d^2
1	24.29	0.00	0.0000
2	24.26	-0.03	0.0009
3	24.17	-0.12	0.0144
4	24.31	0.02	0.0004
5	24.28	-0.01	0.0001
6	24.19	-0.10	0.0100
7	24.33	0.04	0.0016
8	24.50	0.21	0.0441
9	24.30	0.01	0.0001
10	24.23	-0.06	0.0036

$$\langle x \rangle = 242.86/10 = 24.29\,\text{g and } s = \sqrt{(0.0752/9)} = 0.0917,$$

$$\text{range} = \langle x \rangle \pm 2s = 24.29 \pm 0.18\,g$$

or, the test tube weighs between 24.11 and 24.47 g, with 95% certainty.

STANDARD DEVIATION ACTIVITY

Weight of a test tube on 10 different balances

trial	weight	d = X_m - <X>	d^2
1	24.29	0.00	0.0000
2	24.26	-0.03	0.0009
3	24.17	-0.12	0.0144
4	24.31	0.02	0.0004
5	24.28	-0.01	0.0001
6	24.19	-0.10	0.0100
7	24.33	0.04	0.0016
8	24.50	0.21	0.0441
9	24.30	0.01	0.0001
10	24.23	-0.06	0.0036

ACTIVITY: **If each of the values of x_m are checked against the range, note that the weight from balance 8 is outside the range; it should be discarded as unreliable therefore, Using excel, recalculate <x>, d, d^2 and s. Email your excel file to me.**

WORKSHOP 1 precision

1. Bernadette recorded the following data in the lab when determining the percentage of water in an unknown hydrate. What is the percentage water her unknown?

Mass of container:		47.952 g
Mass of container & hydrate:		49.837 g
Mass of container & contents after:		
	First heating:	49.500 g
	Second heating:	48.918 g
	Third heating:	48.811 g

2. An instructor gives a sample of powered metal to each of four students (W, X, Y, & Z), and they weigh the samples on different balances. Their results for three trials are as follows. The true value is 8.720 g.

Student	Trial 1	Trial 2	Trial 3	Average
W	8.72 g	8.7810 g	8.6 g	________
X	8.50 g	8.48 g	8.511 g	________
Y	8.56 g	8.7770 g	8.830 g	________
Z	8.4100 g	8.720 g	8.550 g	________

a) Calculate the average mass for each data set with the correct significant figures.
b) Which student was the most accurate in weighing? ______
c) Which student was the most precise? ______
d) Which student had the best combination for accuracy and precision? ______