Name _____ **College Prep.CHEMISTRY**
Mr. Dove **Based on Chapters 1, 2, and 3 in the textbook**

Unit 1A: Measurement

Objectives:

1. Understand selected metric units and prefixes Convert number to and from scientific notation
2. Measure with significant figures
3. Perform scientific calculations: Percent Error and Density
4. Understand the difference between temperature and heat and Kelvin and Celsius Temperature scales. Convert between the two.

Metric System

- In chemistry (and all sciences) the International System of Units (SI) is used. It is a universal set of units that allows scientists from around the world to be consistent with each other.
- TABLE _____ is a list of _____
- The SI system is a decimal system, meaning prefixes are used to change SI units by a power of 10
- TABLE _____ is a list of _____

Metric Conversion

- Often we will want to convert from one unit of measurement to another. To do so you need a conversion factor.
- Use TABLE _____ to convert from different metric sizes
- **Example**: Convert 55 milligrams (mg) to grams (g):
- Step 1: Create a conversion factor: _____ mg = _____ g
- Step 2: Multiply by your conversion factor. Place the unit you want to convert to in the numerator (top spot) and what you want to cancel in the denominator (bottom spot)
- **Examples**: Complete the following conversions

1) 15 kg to g; _____kg = _____g

2) 500 mg to g; _____mg = _____ g

2) 0.003 m to cm; _____ m = _____ cm

4) 642 cg to kg; _____ cg = _____ g; _____ g = _____ kg

III. Scientific Notation

- Throughout the year we will encounter VERY small and big numbers. We use scientific notation to represent these numbers in powers of tens.
- 5,300,000 m can be written as _____
- 0.00000375 g can be written as _____
- **Examples**: Write the following in scientific notation

1) 34500000 kg = _____

2) 7561000 m = _____

3) 0.000301 cm = _____

4) 0.000000002091 mg = _____

- **Examples**: Convert the following from scientific notation to standard

1) 4.51×10^{3} g = _____

2) 8.91×10^{-4} km = _____

3) 5.12×10^{-6} kg = _____

4) 5.234×10^{7} cm= _____

IV. Significant Figures

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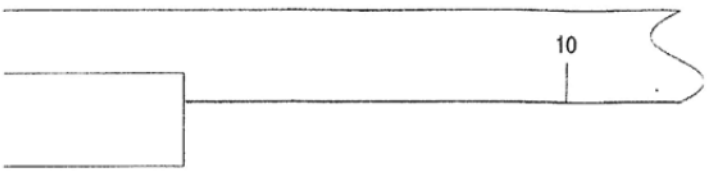
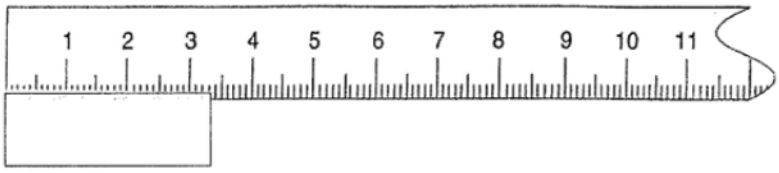
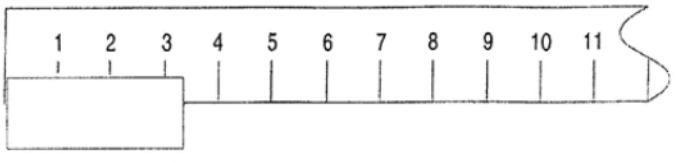


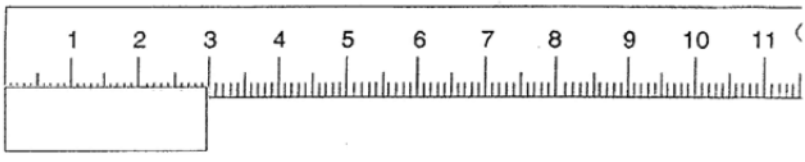
Measurement _____



Measurement _____



Measurement _____



Measurement _____

- **Rules for Counting Significant Figures:**

1. Nonzero digits are always significant	Rule 1 Ex)
2. All final zeros after the decimal point are significant	Rule 2 Ex)
3. Zeros between two other significant digits are always significant	Rule 3 Ex)
4. Zeros used solely as placeholders are not significant	Rule 4 Ex)

- **Examples:** How many significant figures does each of these measurements have?

1) 3.1 m _____

2) 3.0001 kg _____

3) 1.20×10^{-4} km _____

4) 0.007060 cm _____

V. Significant Figures in Calculations

- **Addition/Subtraction**: 10.52 + 349.0 + 8.240

  Step 1:

  Answer: _____

  Step 2:

  New Answer: _____

  Example: 3.21 + 123.1101 + 44.651 = _____ => _____

- **Multiply/Divide**: 7.55 x 0.34

  Step 1:

  Answer: _____

  Step 2:

  New Answer: _____

  Example: 4.501 ÷ 2.10 = _____ => _____

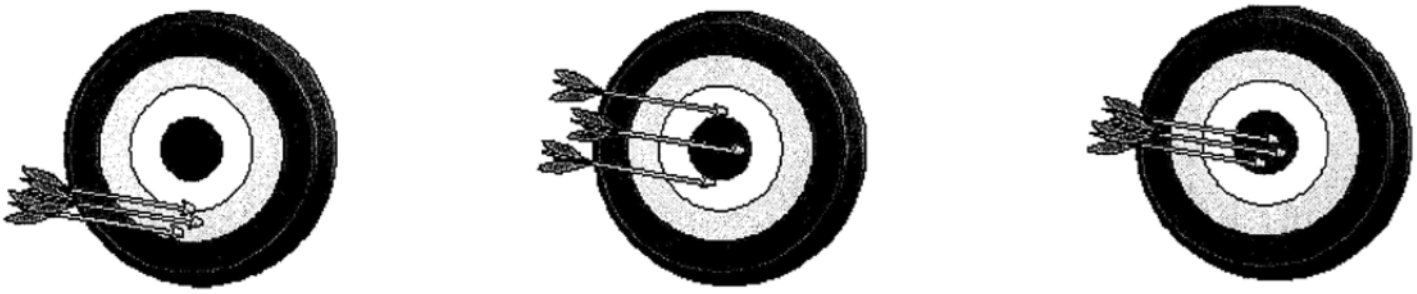
  Note: Constants, conversion factors, or exact numbers (e.g. number of people in a room) are **NOT** taken into account when performing calculations, **ONLY MEASURED QUANTITIES.**

VI. Types of Measurement

- **Quantitative:** _____
- **Qualitative:** _____

VII. Precision vs. Accuracy

- **Precision:** How close a series of measurements are
- **Accuracy:** How close your measurements are to the actual/accepted value



- **IN GENERAL: We want measurements to be accurate (close to accepted) and precise (similar with each other)**

VIII. Percent Error

- **Equation (see reference tables):**

- A student finds the density of copper to be 8.218 g/cm^3. The actual density of copper is 8.960 g/cm^3. Find the percent error in her measurement.

IX. Density

- **Mass:** _____

  Measured in:

- **Volume:** _____

  Measured in:

- **Density:** _____

  Measured in:

- **Equation (see reference tables):**

- **Examples:**

1. A person brings in what he thinks to be a gold ring to a jewelry store. The ring has a mass of 4.5 g and a volume of 0.233 cm^3. Is this a gold ring? (Hint: find the density and compare it on Table S)

   Givens: **Equation and Answer:**

   Want:

2. A piece of scrap metal made of ***iron*** has a volume of 305.5 cm^3. Find the mass of the iron.

   Givens: **Equation and Answer:**

   Want:

3. A rock has a mass of 120.5 g. It is put into 103.35 mL of water and the water rises to 118.42 mL. Find the density of this rock.

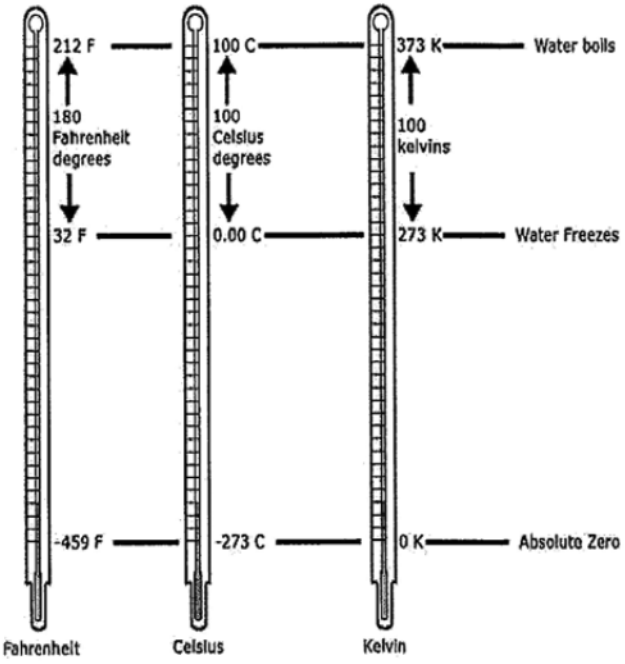
4. What is the density of an unknown block having a mass of 972.3 g and dimensions of 4.57 cm by 15.32 cm by 11.28 cm.

X. Temperature vs. Heat

- **Temperature**:

- **Heat**:

- **Absolute Zero:**



- **Temperature Scales (See Reference Tables):**

  $K = {}^{o}C + 273$ K = Kelvin ${}^{o}C$ = degrees Celsius

- **Convert**:
  A) 200 degrees Celsius to Kelvin B) 500 K to ${}^{o}C$

- **Remember**: Kelvin uses bigger values and is always _____

Unit 1B: Introduction to Chemistry

Objectives:

1. Identify and understand basic chemistry terms: matter, atom, compound, element, mixture
2. Identify and draw particle diagrams for different phases of matter
3. Identify and classify matter as an element, compound, or mixture
4. Draw particle diagrams for each classification of matter
5. Identify between physical and chemical properties/changes
6. Understand mass and matter is conserved in a chemical reaction

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I. Important Terms:

- **Chemistry:** Study of _____
- **Matter:** Any object that has _____

Atom	**vs.**	**Compound**
Definition:		**Definition:**
Particle Diagram		**Particle Diagram**

II. Phases of Matter

- The phase that matter is in depends on:

1.

2.

Phase	Shape	Volume	Particle Diagram	IMF	Movement
Solid (*s*)					
Liquid (*l*)					
Gas (*g*)					

- These are the three most common states and the ones we will focus on for this course. However, there are additional ones such as plasma, supercritical fluid, and degenerate gas.

III. Classification of Matter

- **Element**: Substance that _____

  _____ under normal conditions

- **Compound**: Substance consisting of _____

_____

_____

_____

- **Mixture:** _____

_____

_____

Aqueous Solution (*aq*):

(A)	(B)	(C)
H_2 (g)	H_2O(l)	NaCl (aq)
He (g)	H_2O (g)	$C_6H_{12}O_6$ (aq)
K (s)	NaCl (s)	CO_2 (aq)
Co (s)	CO_2 (g)	Air
Na (s)	$C_6H_{12}O_6$ (s)	Tap water

- **How Elements, Compounds, and Mixtures can be illustrated in Particle Diagrams:**

Element	Compound	Mixture

IV. Physical vs. Chemical Properties

Physical Properties	Chemical Properties
Properties of an element or compound that can be observed or measured _____ _____	The ability of an element or substance to undergo a _____ and form a _____
Examples of Physical Properties:	**Examples of Chemical Properties:**

- **Physical or Chemical Properties**: Determine whether chemical or physical
  1. ___ Water boils at 100 degrees Celsius
  2. ___ Water can be separated by electrolysis into hydrogen and oxygen
  3. ___ Sugar is capable of dissolving in water
  4. ___ Vinegar will react with baking soda
  5. ___ Yeasts acts on sugar to form carbon dioxide and ethanol
  6. ___ Wood is flammable

V. Physical vs. Chemical Changes

Physical Change	Chemical Change
• It does NOT _____, it just changes _____ • A change that does NOT affect a substance's chemical composition	• Changing _____ into a _____ • A color change may occur and a _____
Physical Change Phrases	**Chemical Change Phrases**
Particle Diagram for Physical Change	**Particle Diagram for Chemical Change**

- **Physical or Chemical Change**: Determine whether chemical or physical

1. ____ Dry ice, solid carbon dioxide, is sublimed at room temperature.
2. ____ Iron rusts in damp environment
3. ____ Gasoline burns in the presence of oxygen
4. ____ Hydrogen peroxide decomposes to water and oxygen
5. ____ Burning coal
6. ____ Cooking a steak
7. ____ Cutting grass

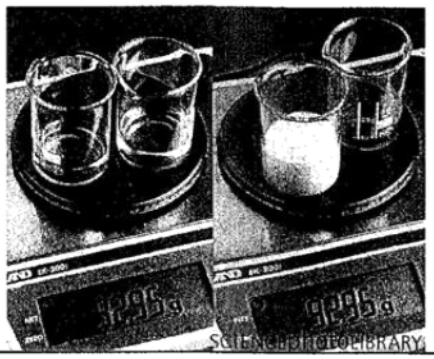
- **Chemical Reaction Equation:**

A chemical reaction ALWAYS results in new substance(s)

$$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$$

VI. Conservation of Mass:

- ***Mass cannot be created or destroyed in a chemical reaction***



Silver nitrate ($AgNO_3$) and sodium chloride (NaCl) solutions before and after chemical reaction

Examples:

1. If 50.0 grams of sodium reacts with chlorine to form 126 grams of sodium chloride. How many grams of chlorine reacted?

2. If 178.8 g of water is separated into hydrogen and oxygen gas, and the hydrogen gas has a mass of 20.0 g. What is the mass of the oxygen gas produced?

VII. Classification of Matter Chart