Unit 1a Matter and Energy

# Matter and Change

## What is Chemistry?

### Chemistry is…

1. The study of matter (**structure, properties**)…
2. The **changes** that matter undergoes…and
3. The **energy** involved in those changes.

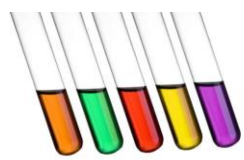
## Learning Targets

1. Classify substances as either **ELEMENTS** or **COMPOUNDS**.
2. Identify the difference between and **PHYSICAL** and **CHEMICAL change**.

## What is Chemistry?

Chemistry involves Chemicals

*Any substance with a definite composition.*



Examples: $H_2O$ and $H_2O_2$

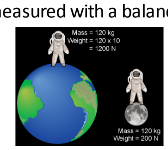
## Matter and Its Properties.

**MATTER:** Anything with mass and volume

**WEIGHT:** versus **MASS:**

*The **force** of gravity on an object.* (measured with a scale)


*The **amount** of matter in an object.* (measured with a balance)




## Matter and Its Properties.

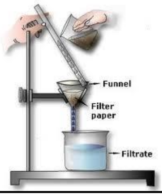
**ATOMS** Smallest unit of matter

| ELEMENT | COMPOUND |
| --- | --- |
| **CANNOT BE** broken down into a simpler substance | **CAN BE** broken down into simpler substances |
| Cu, Fe, Hg, K, Xe | $H_2O$, $H_2SO_4$, $NaCl$, $Fe_2O_3$ |

## Matter and Its Properties.

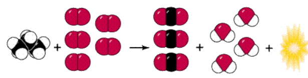
**MIXTURE**

Separated by
**PHYSICAL Process**




**COMPOUND**

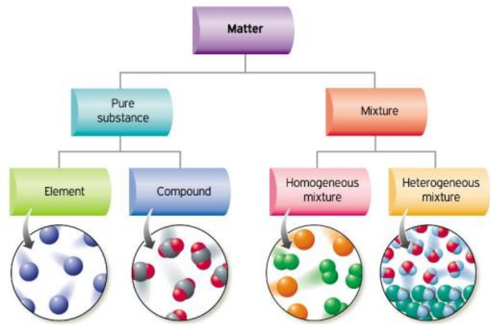
Separated by
**CHEMICAL Reaction**



$C_3H_8$ + 5 $O_2$ → 3 $CO_2$ + 4 $H_2O$ + Heat and light

Propane, Oxygen gas, Carbon dioxide, Water

## Matter and Its Properties.




Physical Property

Measured without changing the substance

- Intensive — **DOES NOT** depend on the amount
- Extensive — **DEPENDS** on the amount

## T-Chart Group Activity

**INTENSIVE** –or – **EXTENSIVE**

- Boiling Point
- Mass
- Volume
- Freezing Point
- Weight
- Energy
- Elasticity
- Color
- Malleability
- Length
- Conductivity
- Luster

Chemical Property

A substances ability to change it into different substances

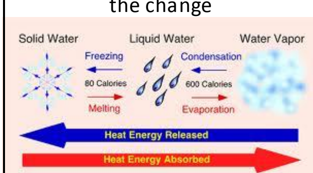
The ability to burn in air
(Coal Burning)

The ability to oxidize
(Iron Rusting)

## Matter and Its Properties.

**PHYSICAL CHANGE**

The same substance
BEFORE and AFTER
the change



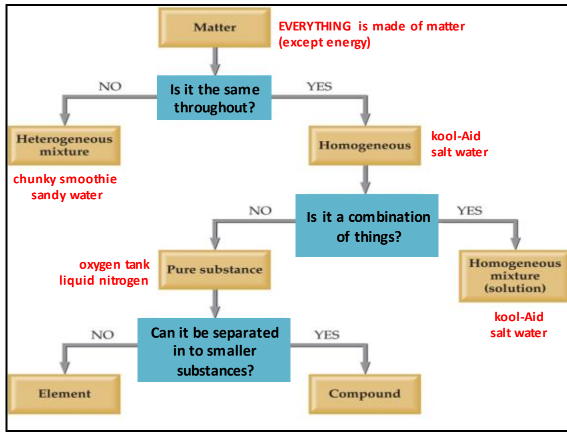

**CHEMICAL CHANGE**

Produces
ONE or MORE
new substances.

## Matter and Its Properties.

**Indicators of a CHEMICAL CHANGE**

1. **Color change**.
2. Gas is given off. **(bubbles)**
3. Precipitate. **(solid formed)**
4. Change in energy. **(HOT** or **COLD)**
5. Light emission. **(Light given off)**




Unit 1b Matter and Energy

## Measurements and Calculations

## Unit 1b Learning Targets

1. Classify substances as either elements or compounds.
2. Tell the difference between physical/chemical properties and changes.
3. Identify proper measurement techniques and possible errors.
4. Understand energy moves between a system and its surroundings.

## Units of Measure.

**Qualitative**
- Descriptive (Non-numeric)
- The fact that the sky is blue

**Quantitative**
- Measurement (Numeric)
- A copper sample has a mass of 25.7 grams

## Units of Measure.

The International System (SI). 1960s

Commonly used SI units in CHEMISTRY.

1. Length meters (m)
2. Mass kilograms (**kg**)
3. Time seconds (s)
4. Volume liter (L)
5. Quantity moles (mol)

## Units of Measure.

- The International System (SI) 1960s
- SI units PREFIXES.

| Prefix | Symbol | How many in a base unit? |
|---|---|---|
| Nano- | n | 1,000,000,000 |
| Micro- | µ | 1,000,000 |
| Milli- | m | 1,000 |
| Centi- | c | 100 |
| Deci- | d | 10 |

**Example** *1 m = 100 cm*
*1 g = 1,000 mg*

| Prefix | Symbol | How many base units? |
|---|---|---|
| Kilo- | k | 1,000 |
| Mega- | M | 1,000,000 |
| Giga- | G | 1,000,000,000 |

**Example** *1 km = 1000 m*
*1 Mg = 1,000,000 g*

## Units of Measure.

Conversion Factor Practice.

**Example 1** – *Convert 22,000 g to kg.*

**Example 2** – *Convert 0.0290 m to mm.*

**Example 3** – *How many kilometers are in 2.34 miles?*

**Example 4** – *How many meters are in 0.62 ft?*

## Metric Conversion Mnemonic

**King Henry Died By Drinking Chunky Milk**

Kilo Hecto Deka BASE Deci Centi Milli

## Units of Measure.

**Conversion Factors.**

- *Math used to relate 2 units that measure the same quantity (written as a fraction);* **Equal to 1.**
- Example 1 m = 1000 mm. $\frac{1\ m}{1,000\ mm}$

*The distance between the North and South Building is roughly 12,672 inches. How many feet do you travel between class?*

## Conversion Group Activity

1. Divide into **PAIRS**.
2. Each PAIR is assigned one **DATA GROUP**.
3. Each PAIR is responsible for organizing the data group from the **SMALLEST to LARGEST measurement**.

## Conversion Group Activity

| Length (White) | Mass (Yellow) | Volume (Green) |
|---|---|---|
| 10 mm | 10 mg | 10 mL |
| 2.5 x $10^{-3}$ hm | 2.5 x $10^{-3}$ hg | 2.5 x $10^{-3}$ hL |
| 1.0 m | 1.0 g | 12.0 fl. oz. |
| 17.2 dm | 17.2 dg | ¼ gallon |
| 6 ft. 1 in. | 1316 cg | 1.0 L |
| 1316 cm | 1.2 oz. | 17.2 dL |
| 0.33 km | ¼ pound | 1316 cL |
| ¼ mile | 0.33 kg | 0.33 kL |

## Thursday Sept 4

**PREPARE YOURSELF!**

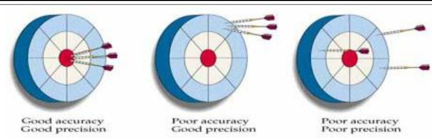
Homework Check over

**UNIT CONVERSIONS!**

Unit 1 Matter and Energy

## Accuracy vs. Precision
## Significant Figures

## Using Scientific Measurement.



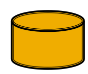

**Accuracy**

How close a set of measurements are to the **CORRECT VALUE.**

**Precision**

How close a set of measurements are **GROUPED** regardless of being correct.

## Accuracy vs. Precision



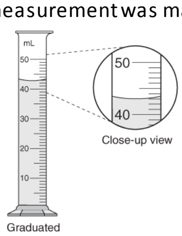
Mass = 13.6 grams

| | Trial 1 | Trial 2 | Trial 3 | Average |
|---|---|---|---|---|
| Student 1 | **13.5** | | | |
| Student 2 | **13.6** | | | |
| Student 3 | **9.4** | | | |

## Using Scientific Measurement.

**Significant Figures.**

All digits that occupy places for which **ACTUAL** measurement was made + the last estimated digit.




Significant means "**important**", so we are looking for all of the numbers that show the **ACCURACY** within each measurement

## Significant Figures

How many sig figs? State the rule(s) [Proof]

1. 1234 kg
2. 0.023 L
3. 9010.0 mm
4. 0.0001 g
5. 1078.0010 mL
6. 1,020,010 km

## Significant Figures

How many sig figs?

1. 1234
2. 0.023
3. 9010.0
4. 0.0001
5. 1078.0010
6. 1,020,010

Sig fig rounding.

30.495 g
-15.60 g
**14.895** ⇒ **14.90 g**

$\frac{14.90\ g}{3.1\ mL}$ **= 4.806451613** ⇒ **4.8 g/mL**

## Using Scientific Measurement.

**Scientific Notation.**

*How scientists show either BIG or SMALL numbers.*

- 602,200,000,000,000,000,000,000 = 6.022 x $10^{23}$

The decimal point is always located between the first and second digit AND the first digit must be non-zero number.

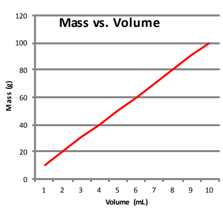
65000000. (7 6 5 4 3 2 1) → 6.5x$10^{7}$

.0000987 (-1 -2 -3 -4 -5) → 9.87x$10^{-5}$

## Using Scientific Measurement.

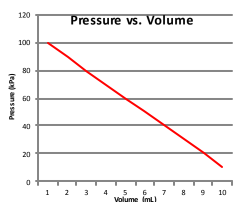
**DIRECT PROPORTION**

*As one variable increase, so does the other.*




**INVERSE PROPORTION**

*As one variable increases, the other decreases or vice versa*




## Using Scientific Measurement.

Experiments will always have errors.
(human, mechanical, environmental)

**PERCENT ERROR** (the lower the percent, the better)

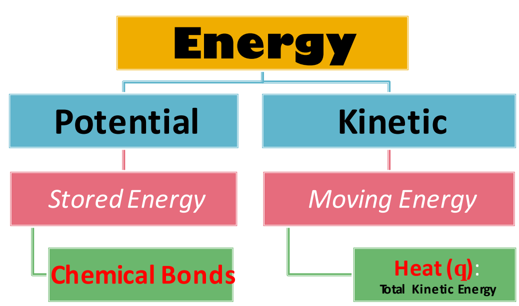
*Determines the accuracy of the experiment.*

$$Percent\ Error = \frac{|\text{Lab Measurement} - \text{TRUE Value}|}{\text{TRUE Value}} \times 100$$

Unit 1 Matter and Energy

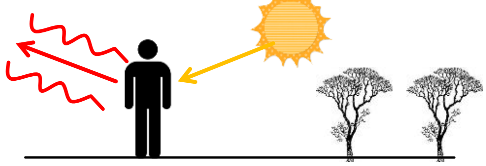
## Energy Transfer

## What is ENERGY?



## Energy Transfer.

**HEAT** Energy moves back and forth between a system and its surroundings.



**Temperature:** *Average KE of the random motion of particles in a substance.*

## Energy Units.

| Name |
|---|
| Joule (J)<br>SI Unit |
| Calorie (C)<br>1 Calorie = 4184 J |
| 140 C = 585,760 J |


## Energy Transfer.

**Specific Heat Capacity ($C_p$)**

*The ability for a substance to absorb heat.*

| LOW vs HIGH specific heat capacity | |
|---|---|
| **Metals** | **Nonmetals** |
| Great Conductors | Poor Conductors |
| **LOW** heat capacity (Less than 1.0) | **HIGH** heat capacity (Greater than 1.0) |

## Energy Transfer.

$$q = (m)(\Delta T)(C_p)$$

| | | |
|---|---|---|
| q | = | Heat (**J**) |
| m | = | Mass (**g**) |
| $\Delta T$ | = | Change in Temperature (**$\Delta T = T_f - T_i$**) |
| $C_p$ | = | Specific Heat Capacity (**Given**) |

## Practice 1

Calculate the amount of heat (in Joules) needed to raise 34 g $H_2O$ from 55°C to 67°C.

**$q = (m)(\Delta T)(C_p)$**

$C_p$ $H_2O$ = 4.184 J/g°C

q = (34 g) ([67 – 55 °C]) (4.184)

q = (34 g) (12 °C) (4.184)

q = **1707.07 J**

## Practice 2

Calculate the amount of heat (in Joules) needed to raise 65 g copper from 30°C to 95°C.

**$q = (m)(\Delta T)(C_p)$**

$C_p$ Cu = 0.385 J/g°C

q = (65 g) ([95 – 30 °C]) (0.385)

q = (65 g) (65 °C) (0.385)

q = **1626.63 J**

## $q = (m)(\Delta T)(C_p)$

*The specific heat of water is* ***4.184*** J/g°C.

1. 40.0 g of water is heated from 10.0°C to 30.0°C.
2. 135.6 g of water is cooled from 95.8°C to 21.6°C.
3. 30.0 g of aluminum is heated from 15.0°C to 35.0°C.
4. 450.0 g of iron is cooled from 125.0°C to 45.0°C.
5. 62.3 g of lead is heated from 21.7°C to 136.4°C.

## $q = (m)(\Delta T)(C_p)$

*The specific heat of water is* ***4.184 J/gK****.*

1. ______________
2. ______________
3. ______________
4. ______________
5. ______________

| Specific Heat of Common Substances | |
|---|---|
| Water (l) | 4.184 J/(g•K) |
| Aluminum | 0.897 J/(g•K) |
| Iron | 0.449 J/(g•K) |
| Lead | 0.129 J/(g•K) |