Unit 1: CHEMISTRY AS A SCIENCE

Chapter 1: The Science of Chemistry

1.1: What is Chemistry?

Chemistry: - an organized study of matter and energy.

Chemical: - a substance that has a definite composition – always made of the same matter regardless where the chemical comes from.
- can be natural or manufactured (synthetic).

Chemical Reaction: - a process where chemicals are added or change into other kinds of chemicals.

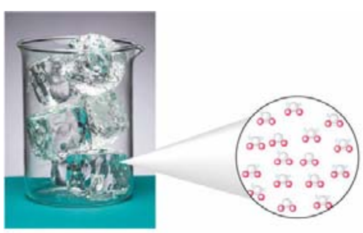
Physical Change: - a change in matter that does **not** alter the composition or identity of substance.

Kinetic Molecular Theory and States of Matter: - particles in matter are constantly in motion.

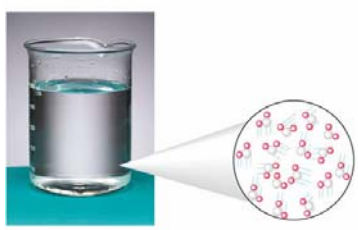
Solid: - the state of matter where it has a definite volume with a constant shape.

Liquid: - the state of matter where it has a definite volume but an indefinite shape.

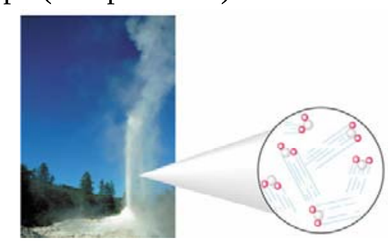
Gas: - the state of matter where it has an indefinite volume and shape (compressible).



Solids have particles are in fixed positions.

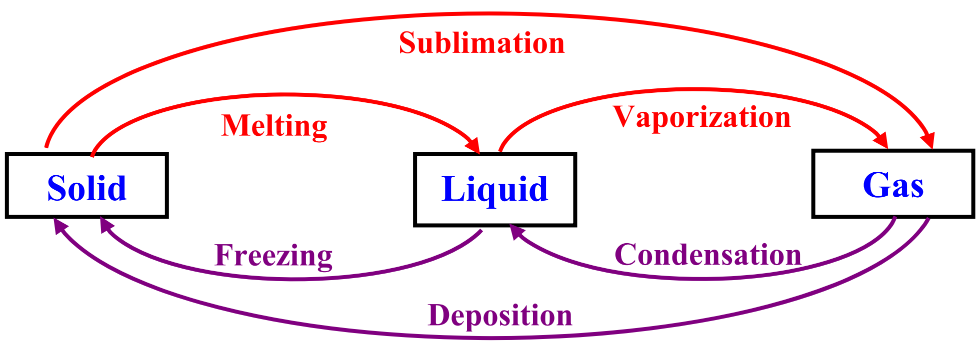


Liquids have particles that can "rolled" past one another.



Gases have particles that have truly random motion and have very weak interactions between molecules.

Phase Change: -when matter changes from one phase into another.



Sublimation: - the phase change from solid to gas directly (vice versa is called **deposition**).

Example: Dry ice (Solid Carbon Dioxide) sublimates from solid to gas directly, skipping the liquid phase.

Chemical Change: - a change in matter that involves altering the composition or identity of substance.
- a change where **New Substance(s)** are formed.

Reactants: - chemicals that go into a reaction.

Products: - chemicals that are produced from a reaction.

$$\textbf{Reactants} \xrightarrow{\text{“yields”}} \textbf{Products}$$

Example: Hydrogen + Oxygen → Water
$H_{2\,(g)}$ + $O_{2\,(g)}$ → $H_2O_{(l)}$

Evidences of a Chemical Change: (*For a new pure substance formed*)

1. **Precipitate (New Solid) ↓ is formed.**
2. **Colour Change.**
3. **Presence of Bubbles or New Odour to indicate a New Gas ↑.**
4. **Heat is suddenly Given off or Taken in or an Explosion!**

Assignment
1.1 pg. 9 #3 to 14

1.2: Describing Matter

Matter: - anything that has a mass and occupies space.

Volume: - an amount of space an object occupies

Mass: - the amount of stuff in an object.

Weight: - the amount of gravitational force that is pulling on an object.

Example: An object that has 50.0 kg on Earth will have a mass of 50.0 kg on the moon. However, the same object, which has a weight of 490.5 N on Earth, will only weight 81.75 N on the moon. This is because the gravitation pull on the moon is one-sixth of that on Earth.

Quantity: - the numerical part of a measurement.

Unit: - the standard part of the measurement.

Example: An object has a mass of 50.0 kg. (50.0 is the quantity and kg is the unit.)

SI Units: - International Metric Units (*le Système International*).

Metric Prefixes and Exponential Notations:

Giga	* *	Mega	* *	Kilo	hecto	deca	Basic Units	deci	centi	milli	* *	micro	* *	nano
G		M		K	h	da	metre (m)	d	c	m		μ		n
10^9		10^6		10^3	10^2	10^1	Litre (L)	10^{-1}	10^{-2}	10^{-3}		10^{-6}		10^{-9}
							gram (g)							
							Kelvin (K)							
							Pascal (Pa)							
							Newton (N)							
							Mole (mol)							

$\times 1000$ **1 mL = 1 cm^3** $\times 1000$

1 L = 1000 cm^3

Note: **1 m^3 ≠ 1 L**

1 m^3 = 1 m × 1 m × 1 m
1 m^3 = 100 cm × 100 cm × 100 cm
1 m^3 = 1,000,000 cm^3
1 m^3 = 1000 L

Example 1: Complete the following unit conversions

a. 345 mL = **0.345** L **(left 3 places)**

b. 42 g = **0.042** kg **(left 3 places)**

c. 54300. m = **54.300** km **(left 3 places)**

d. 26 μm = **0.000 002 6** dam **(left 7 places)**

e. 1854 cm = **0.01854** km **(left 5 places)**

f. 0.035 000 kg = **35000.** mg **(right 6 places)**

USCS: - United States Customary Systems, commonly known as the imperial system.

Length			
1 mile = 1760 yards (yd)	1 foot = 12 inches	1 yard = 0.9144 m	1 inch = 2.54 cm
1 yard = 3 feet	1 mile = 1.609 km	1 foot = 0.3048 m	
Mass		**Volume**	
1 ton = 2000 pounds (lbs)	1 pound = 0.454 kg	1 gallon = 4 quarts	1 gallon = 4.546 L
1 pound = 16 ounces	1 ounce = 28.35 g	1 quart = 2 pints	1 US gallon = 3.785 L

Conversion Factor: - a simple ratio that relates two units that express a measurement of the same quantity.

Example 2: Convert the following units.

a. 15.2 ft = ________________ yd

$$15.2\ \cancel{\text{ft}} \times \frac{1\ \text{yd}}{3\ \cancel{\text{ft}}} = 5.07\ \text{yd}$$

15.2 ft = **5.07 yd**

b. 32.7 ton = __________________ lbs

$$32.7\ \cancel{\text{ton}} \times \frac{2000\ \text{lbs}}{1\ \cancel{\text{ton}}} = 65{,}400\ \text{lbs}$$

32.7 ton = **65,400 lbs**

c. 50. pints = ________________ quarts

$$50\ \cancel{\text{pints}} \times \frac{1\ \text{quart}}{2\ \cancel{\text{pints}}} = 25\ \text{quarts}$$

50 pints = **25 quarts**

d. 5 ft 10. inches = ______________cm

5 ft 10 inches = (5 × 12 + 10) inches = **70 in**

$$70\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 177.8\ \text{cm}$$

5 ft 10 inches = **178 cm**

e. 635 m = ________________________ yd

$$635\ \cancel{\text{m}} \times \frac{1\ \text{yd}}{0.9144\ \cancel{\text{m}}} = 694.4\ \text{yd}$$

635 m = **694 yd**

f. 1024 lbs = _______________________ kg

$$1024\ \cancel{\text{lbs}} \times \frac{0.454\ \text{kg}}{1\ \cancel{\text{lb}}} = 464.9\ \text{lbs}$$

1024 lbs = **464.9 kg**

g. 46.1 L = ____________________ US gallon

$$46.1\ \cancel{\text{L}} \times \frac{1\ \text{US gallon}}{3.785\ \cancel{\text{L}}} = 12.18\ \text{US gallon}$$

46.1 L = **12.2 US gallon**

Derived Units: - units that are derived by multiplying or dividing the base units

Examples: Volume (length × width × height = cm × cm × cm = cm^3);
Speed (metre per second – m/s); Density (mass per volume – g/mL);

Physical Property: - an observation or measurement that does **not** change the composition or identity of substance.

Density: - the amount of mass per unit of volume

$$\textbf{Density} = \frac{\textbf{Mass (g or kg)}}{\textbf{Volume}\ (\mathbf{cm^3, mL, L, m^3})} \qquad D = \frac{m}{V}$$

Example 3: Lead has a density of 11.34 g/cm^3. If a lead sphere has a radius of 5.00 cm, what is its mass?

D = 11.34 g/cm^3 r = 5.00 cm m = ?

Manipulate the formula to solve for m:

$$D = \frac{m}{V}$$

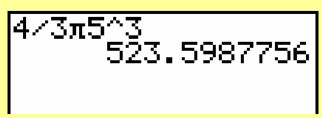
$$DV = m$$

We need to use the Volume formula of a Sphere.

$$V_{sphere} = \frac{4}{3}\pi r^3$$

$$V_{sphere} = \frac{4}{3}\pi(5.00\ \text{cm})^3$$

$$V_{sphere} = 523.5987\ldots\ \text{cm}^3$$



(Do NOT round off. We are not done yet.)

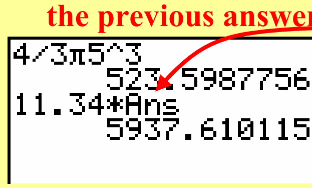
Substitute D and V to solve for m

To recall all digits of the previous answer

$$m = DV$$

$$m = (11.34\ \text{g/cm}^3)(523.5987\ldots\ \text{cm}^3)$$

$$m = 5.94 \times 10^3\ \text{g or } 5.94\ \text{kg}$$



2nd **ANS** **(−)**

Chemical Property: - an observation or measurement that involves a change in the composition or identity of substance.

Examples: Combustability (ease of combustion) and reaction rate

Assignment

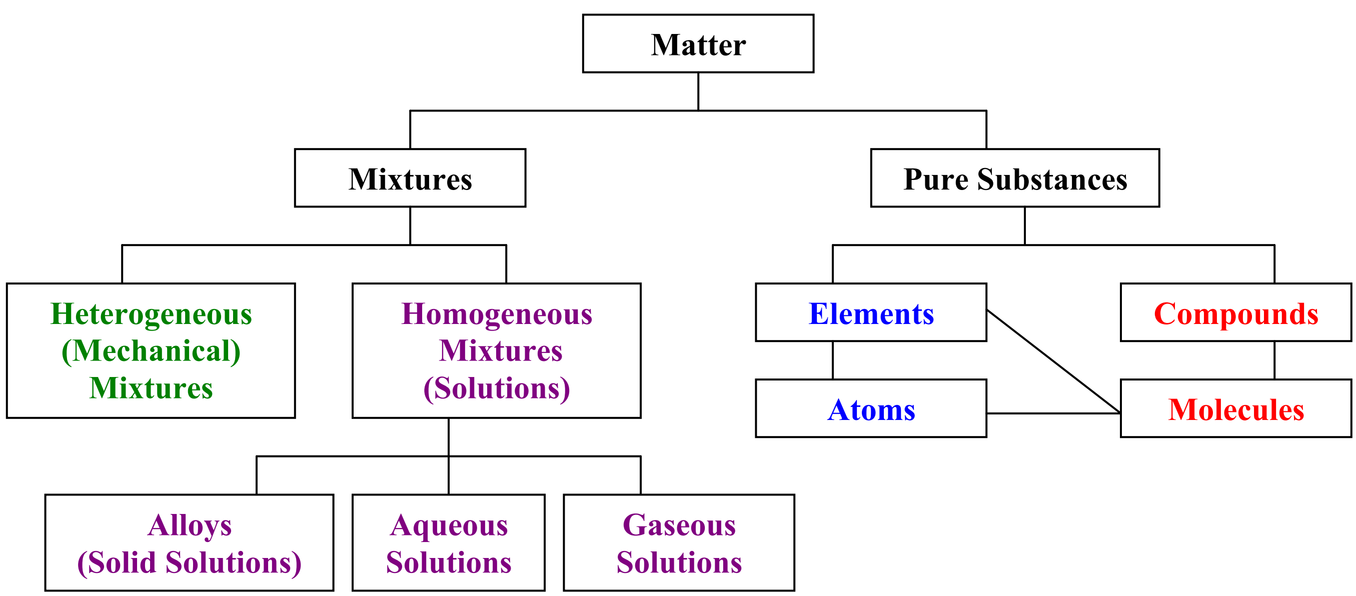
1.2 pg. 14 #1 to 3; pg. 19 #1 to 14

1.3: Classification of Matter

All substance in the universe is made up of **matter**. However, matter can be classified into **mixtures** and **pure substances**.

There are two kinds of mixtures. **Heterogeneous (*hetero* means different) mixtures** are mixtures which we can see its different components with the naked eye (also called **mechanical mixtures**). An example of a heterogeneous mixture is a bag of assorted nuts. We can clearly see the different kind of nuts (walnuts, peanuts, chestnuts, hazelnuts ... etc.) in the bag. A **homogeneous (*homo* means the same) mixture** is also called a **solution**. Unlike heterogeneous mixture, a solution is a mixture that consists of different components, which cannot be seen from a naked eye. An example of a solution is a salt solution. After we completely dissolved the salt in water, we cannot see the salt particles in the water.

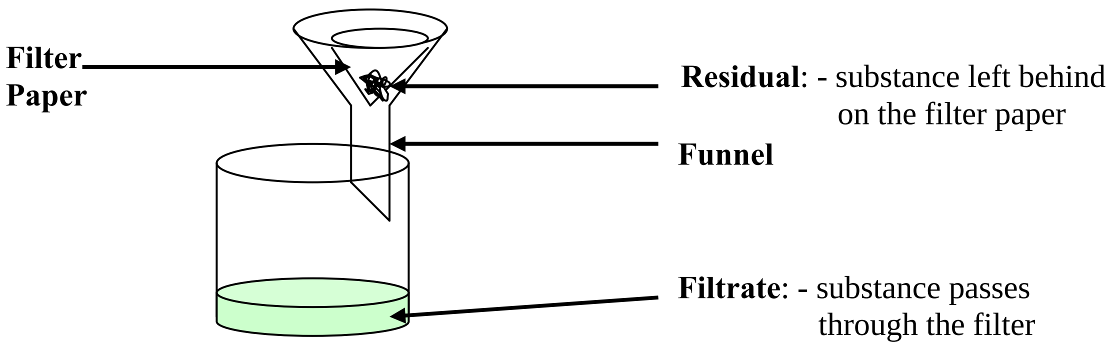
Unlike mixtures, **pure substance** is a substance with a constant composition that cannot be separated by physical means (like phase changes and temperature changes). Pure Substances can be classified into **elements** and **compounds**. Element is a pure substance that has one kind of **atom**. The Periodic Table of Elements lists all the different elements that are either found in nature or prepared in the laboratory synthetically. An atom is defined as the smallest particle of matter. An example of an element is hydrogen. It contains only hydrogen atoms. A compound is defined as a pure substance that is composed of two or more different elements. The smallest unit of a compound is called a **molecule** (a particle that is made up of two or more different atoms or a **unit** of two or more identical atoms). An example of a compound is water. The smallest unit of water is the H_2O molecule. Each water molecule (H_2O) contains two hydrogen atoms and an oxygen atom. An element can have molecular units. An example of that is hydrogen. In its natural state, hydrogen gas exists as H_2 molecules, which consist of units of two hydrogen atoms. **Other diatomic elements (elements that have pair of atoms as molecules) besides H_2 are O_2, N_2, F_2, Cl_2, Br_2 and I_2.** Most of the rest of the elements exist as **singular atomic units (monoatomic)**. Iron atoms, for example, do not organize themselves in multiple atomic units. They exist as individual atoms.



Mixtures: - are matters that are made up of more than one kind of substances and the components can be separated by **Physical Change – No New Substance is formed (change of state, stirring, filtering… etc)**.

Heterogeneous (Mechanical) Mixture: - mixture that is composed of two of more substances where the components can be seen by the naked eye.

Filtration: - using a filter and a funnel, a mechanical mixture consists of liquids and solids can be separated.



Homogeneous Mixture (Solution): - mixture that is composed of two or more substances where the components the same throughout (cannot separate the components by the naked eye).

Solute: - the substance that is being dissolved. **Solvent**: - the substance doing the dissolving

Example: Salt Water (Solute = Salt; Solvent = Water) 9% Alcohol (Solute = Alcohol; Solvent = Water)

Evaporation: - an aqueous solution that consists of a solid solute can be recovered by evaporation of the solvent. The solvent may be recovered as well if a condensation device is used.

Distillation: - an aqueous solution that consists of a liquid solute can be separated by evaporation of the substance with a lower boiling point followed by condensation.

Pure Substance: - a substance with a constant composition.
- in a case where the pure substance is composed of more than one kind of matter, they can only be separated by **chemical change (burning, oxidation, electrolysis … etc)**.

Element: - a pure substance that is made up of one kind of atom.

Compound: - a pure substance that is made up of more than one kind of element.

Atom: - the smallest particle of matter.

Molecule: - the smallest unit of a compound or a diatomic or a polyatomic element.
- basically, it is a particle unit that is made up of more than one atom.

Examples:

1. Classify the following as Heterogeneous or Homogeneous Mixture:
   a) a bag of gravel b) cement c) saturated salt water d) a methanol and water mixture
   e) oil and water f) the atmosphere g) Jell-O h) diet carbonated soft drink

2. Classify the following as Mixture or Pure Substance:
   a) lake water b) tap water c) distilled water d) iron e) steel (iron and carbon)
   f) chromium g) beer h) sugar i) gasoline

3. Classify the following as Element or Compound: (use the Periodic Table of Elements)

a) hydrogen	b) water	c) ammonia	d) oxygen	e) carbon dioxide	f) chlorine
g) ethanol	h) charcoal (carbon)	i) salt	j) nickel	k) gold	l) neon
m) propane	n) baking soda	o) uranium	p) mercury		

Answers:

1. Heterogeneous Mixtures: a), e) Homogeneous Mixtures: b), c), d), f), g), h)

All the components of the heterogeneous mixtures can be seen by the naked eye. However, the components of the homogeneous mixtures cannot be distinguished by the naked eye.

2. Mixtures: a) lake water: contains water, soil particles, micro-organisms ...etc.
b) tap water: contains fluoride and chloride additives.
e) steel: a mixture of iron and carbon. g) beer: contains alcohol, water and other ingredients.
i) gasoline: contains mostly octane and other hydrocarbons.

Pure Substances: c) distilled water: contains water (H_2O) only.
d) iron: an element with a symbol Fe.
f) chromium: an element with a symbol Cr.
h) sugar: a compound commonly known as sucrose ($C_{12}H_{22}O_{11}$).

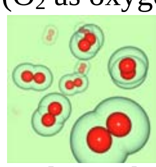
3. Elements: a) hydrogen (H) d) oxygen (O) f) chlorine (Cl)
h) carbon (C) j) nickel (Ni) k) gold (Au)
l) neon (Ne) o) uranium (U) p) mercury (Hg)

Compounds: b) water (H_2O) c) ammonia (NH_3) e) carbon dioxide (CO_2)
g) ethanol (C_2H_5OH) i) salt (NaCl) m) propane (C_3H_8)
n) baking soda ($NaHCO_3$)

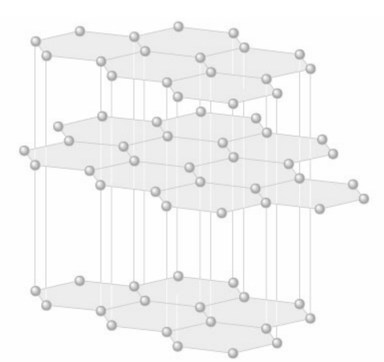
*If the name of the substance appears on the Periodic Table of Elements, then it is an element.

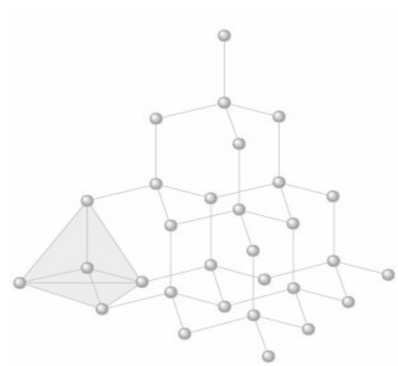
Allotrope: - one of the different molecular form of an element.

Examples: Oxygen can be diatomic or triatomic (O_2 as oxygen gas and O_3 as ozone gas).



Carbon can be graphite (layers), diamond (tetrahedral), or as C_{60} (Bucky Ball).





Assignment

1.3 pg. 28 #1 to 9, 11 to 14; pg. 31−32 #11, 13, 14, 16 to 23, 26 to 28

Chapter 2: Matter and Energy

2.1: Energy

Energy: - is the capacity or the ability to do work (force applied in the same direction over some distance).
- the common unit for energy is **Joule (J)**.
- the USCS unit for energy is **calorie (1 cal = 4.184 J)**.
- energy in food items use the unit **Calorie** **(1 Calorie = 1 *kilo*calorie = 1000 cal = 4184 J).**

Potential Energy: - stored energy.
- **Gravitation Potential Energy**: - stored energy due to an object's position.
- **Elastic Potential Energy**: - stored energy as the elastic is stretched.
- **Chemical Potential Energy**: - stored energy within atoms and molecules of a chemical.

Kinetic Energy: - energy involved when an object is in motion.
- when potential energy is transformed into kinetic energy, work is done.

Example: The water in the reservoir (gravitational potential energy) falls through the dam (kinetic energy) to turn the electric turbine (mechanical energy) to generate electricity.

Physical Change: - a change that only affects the physical properties of matter.
- no new substance is formed, and the process is easily reversed.

Examples: Phase changes, Temperature change, Cutting a substance into smaller bits, Making a Physical Mixture (either Mechanical Mixture or a Solution).

Chemical Change: - a change that only affects the chemical properties of matter.
- whenever a new substance is formed, and the process is **NOT** easily reversed.

Examples: General Chemical Reactions, Rusting, Rotting or Decomposition, Burning.

Work (*w*): - when force is applied over a displacement in the same direction ($w = F \times d$).

Heat (*q*): - the transfer of energy between two objects (internal versus surroundings) due to the difference in temperature.

Temperature: - the average kinetic energy of a substance.

$$T_F = \frac{9}{5}T_C + 32$$

Convert from degree Celsius to Fahrenheit

$$T_F - 32 = \frac{9}{5}T_C$$

32 is an exact number

$$\frac{5}{9}(T_F - 32) = T_C$$

Convert from Fahrenheit to degree Celsius

Some important temperatures:

0°C = Water Freezes
100°C = Water Boils

37°C = Normal Body Temperature
20°C = Ambient Room Temperature

Kelvin: - temperature scale where **0 K (absolute zero) = –273.15°C** (freezing point of hydrogen – no heat, particles stop moving)

$$\boxed{T_K = T_C + 273.15}$$

Example 1: With wind chill, Calgary can get down to −37.0°C. Convert the temperature to Fahrenheit and Kelvin.

$$T_F = \frac{9}{5}T_C + 32 \qquad T_K = T_C + 273.15$$

$$T_F = \frac{9}{5}(-37.0) + 32 \qquad T_K = -37.0 + 273.15$$

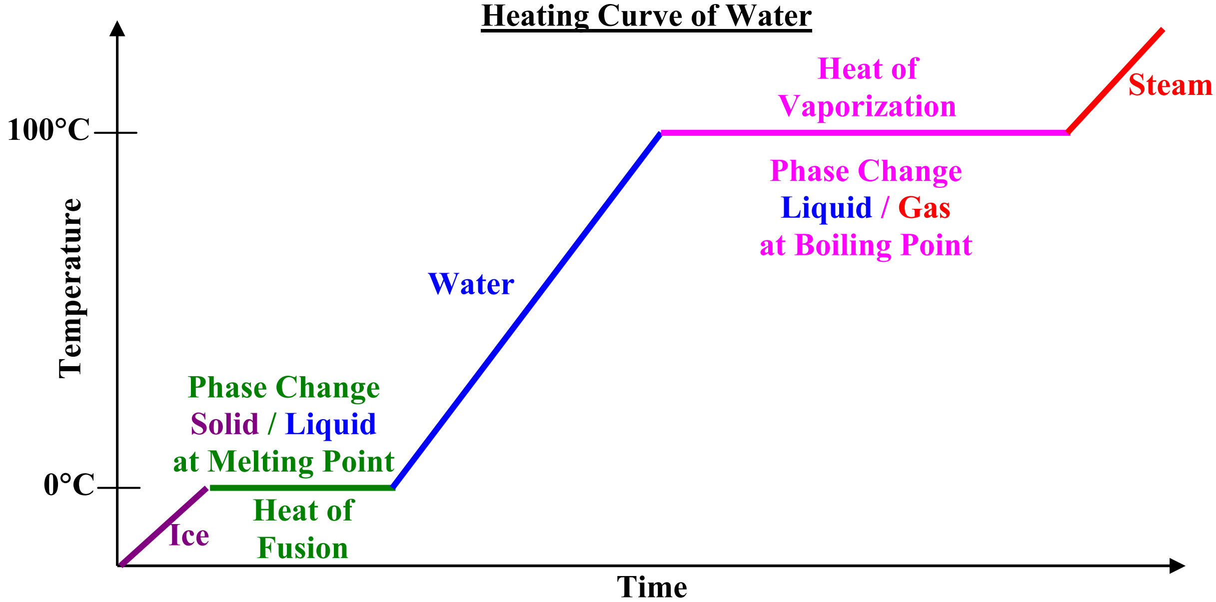
$$T_F = -34.6\ \text{F} \qquad T_K = 236.15\ \text{K}$$

$$T_F = -34.6\ \text{F} \qquad T_K = 236.2\ \text{K}$$

Energy involved in Physical Change (Temperature or Phase Change):

Heating Curve: - a graph of temeparture versus time as a substance is heated from a solid phase to a gaseous phase.
- when a substance is undergoing a **phase change**, its **temperature remains at a constant (the plateau on the heating curve)** until all molecules aquired enough energy to overcome the intermoelcular forces necessary. This is commonly refered to as the **potential change** of a substance.
- when a substance is undergoing **temperature change** within a particular phase, it is refered to as **kinetic change** (because temeperature is also refered to as the average kinetic energy of a substance).

Heating Curve of Water



Specific Heat (c_p): - the amount of heat (J or kJ) needed to change (1 g or 1 kg) of substance by 1°C or 1 K.
- the stronger the intermolecular forces, the higher the specific heat capacity.

System: - a part of the entire universe as defined by the problem.

Surrounding: - the part of the universe outside the defined system.

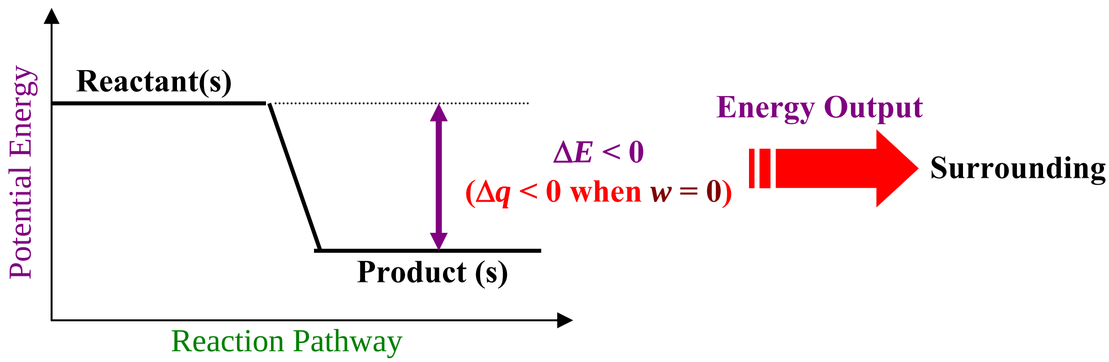
Open System: - a system where ***mass and energy* can interchange** freely with its surrounding.

Closed System: - a system where ***only energy* can interchange freely** with its surrounding but mass not allowed to enter or escaped the system.

Isolated System: - a system ***mass and energy* cannot interchange** freely with its surrounding.

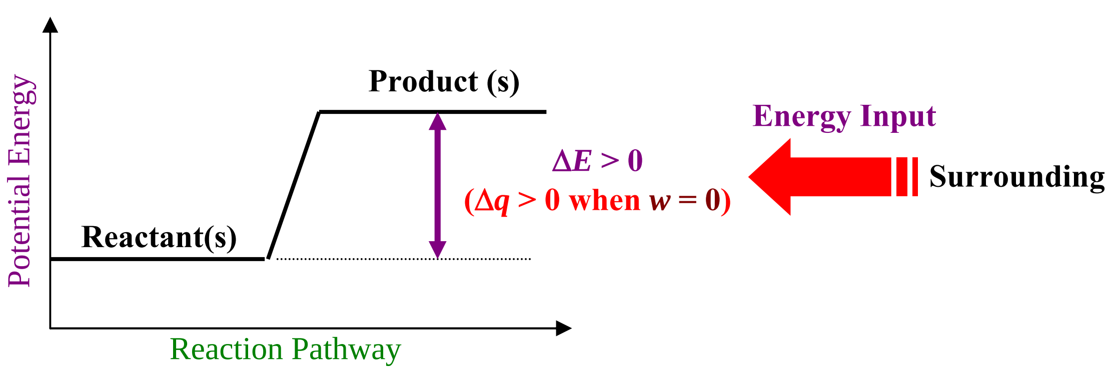
Exothermic Process (**$\Delta E < 0$**): - when **energy flows "out" of the system** into the surrounding.
(Surrounding gets Warmer.)

Potential Energy Diagram for Exothermic Process



Endothermic Process (**$\Delta E > 0$**): - when **energy flows into the system** from the surrounding.
(Surrounding gets Colder.)

Potential Energy Diagram for Endothermic Process



First Law of Thermodynamics: - states that energy cannot be created or destroyed. It can only be converted from one form to another. Therefore, energy in the universe is a constant.
- also known as the **Law of Conservation of Energy** (**$\Sigma E_{\text{initial}} = \Sigma E_{\text{final}}$**).

Assignment
2.1 pg. 45 #1 to 3, 5, 7 to 13

2.2: Studying Matter and Energy

Scientific Method: - a logical method to find solutions of scientific problems.

Steps of Scientific Method

1. **Observation**: - an act of recognizing and noting a fact or occurrence.
2. **Scientific Hypothesis**: - an **educated guess** or a ***testable*** **assumption** to explain any observable phenomenon.
3. **Experimentation** or **Control Test**: - a test performed by scientist and researcher to increase the accuracy and reliability of an experimental test.
   - involves testing of two variables **(manipulated and responding variables)** while all other variables are controlled.

   a. **Accuracy**: - sometimes refer to as **validity**. It describes whether the result is correct.

   b. **Reproducibility**: - sometimes refer to as **reliability**. It describes the consistency and the repeatability of the result.

   c. **Observations**: - often involve making measurements with scientific instrument(s).
4. **Theory**: - an idea that can explain a set of observations that has stood up to repeated scrutiny.
5. **Scientific Law**: - a concise statement that summaries the results of many observations and experiments.
   - describes the phenomenon without trying to explain it.

Limitations of Science: - science cannot answer all questions. It can only tackle "testable" hypothesis.
Philosophical and Religious Questions CANNOT be answered by science.

Law of Conservation of Mass: - mass is neither created nor destroyed in a chemical reaction. (*Lavoisier*)

Example 1: A 3.50 g of iron nail is allowed to rust. The rusted nail has a mass of 5.1 g. What is the amount of oxygen reacted with the iron nail?

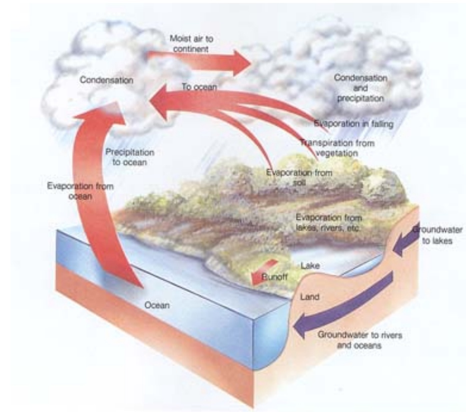
$$\underset{3.50\text{ g}}{\text{Iron}} + \underset{?\text{ g}}{\text{Oxygen}} \rightarrow \underset{5.1\text{ g}}{\text{Rust}}$$

Mass of Oxygen = 5.1 g – 3.50 g

Mass of Oxygen = 1.6 g

Physical Model: - a model that represents the object in a scale that can easily be understood.

Examples: Physical Models of the Solar System and a Animal Cell



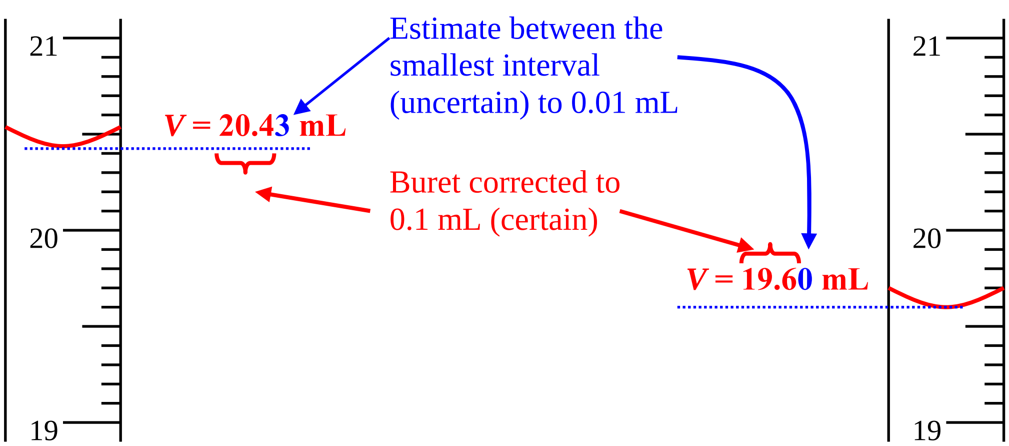
Conceptual Model: - a representation of a *system* in order to describe how components of this system behaves and relates to one another.
- the more accurate a conceptual model, the better it is to predicts the interactions between these components.

Examples: Earth's Hydro Cycle and as systems with representations of how their components interact.

Assignment
2.2 pg. 53 #1 to 13

2.3: Measurements and Calculations in Chemistry

Uncertainty: - all measuring instruments have some levels of uncertainty due to how they were manufactured or users' reading errors.



Precision: - the smallest interval of a measuring instrument.
- uncertainty is sometimes calculated by half the precision of an electronic instrument.

Example: An electronic balance with a precision of 0.01 g has an uncertainty of ±0.005 g

$$\textbf{Uncertainty} = \pm \frac{\textbf{Precision}}{\textbf{2}}$$
(electronic instrument)

Scientific Notation: - commonly used to state very big or very small numbers.

$(1 \text{ to } 9.999\ldots) \times 10^n$

***n* is an integer**
If $n < 0$, then the actual number was smaller than 1
If $n > 0$, then the actual number was greater than 10

Example 1: Convert the following standard notations to scientific notations or vice versa.

a. Speed of Light = 3×10^5 km/s = **300,000 km/s** **(moved 5 decimal places to the right)**

b. Mass of an Electron = 9.11×10^{-31} kg = **0.000 000 000 000 000 000 000 000 000 000 911 kg** **(moved 31 decimal places to the left)**

c. Diameter of a Red Blood Cell = 0.000 007 5 m = **7.5×10^{-6} m** **(moved 6 decimal places to the right)**

d. 2003 US Debt = $6,804,000,000,000 = **$$6.804 \times 10^{12}$** **(moved 12 decimal places to the left)**

Exact Number: - number that indicates no uncertainty. (Numbers in formulas; numbers written in words, counting numbers, or container size)

Examples: **Two** chairs, $SA = 4\pi r^2$, **2500** atoms, **100** mL Beaker

Significant Digits (Figures): - digits used in the measurement plus one uncertain value (*for non-electronic measuring devices* like graduated cylinder, ruler … etc.).

To Count Significant Digits

1. Start counting the first non-zero digit. Do NOT count the leading zero(s).
2. Count all **captive zeros** (between non-zero digits) and **trailing zero** at the end of the measurement.
3. Include ALL digits of a whole number ***if*** it does not contain a decimal point.
   (Examples: 420 → 3 sig digs; 402 → 3 digits)

Example 2: State the number of significant digits for the following measurements.

a. 0.03 g	1 significant digit	e. 25 000 g	5 significant digits
b. 0.030 g	2 significant digits	f. 9.300×10^4 m	4 significant digits
c. 0.0304 g	3 significant digits	g. 4.05×10^{-2} L	3 significant digits
d. 0.03040 g	4 significant digits	h. 7000 °C	4 significant digits

Calculating with Significant Digit

1. **Adding and Subtracting**: - Line up the significant digits. The answer should be to the **least precise measurement** used in the calculation.

Example 3: 5.345 g + 0.42 g + 11.8 g

```
   5.345
   0.42
 +11.8
 -------
  17.565
```

Least precise decimal place is at the tenth. Final Answer should be to one decimal place.

17.6 g

2. **Multiplying and Dividing**: - answer should be in the least number of significant digits used in calculation.

Example 4: $\dfrac{13.25\text{ g}}{1.02\text{ mL}}$

$\dfrac{13.25\text{ g}}{1.02\text{ mL}} = 12.99019608$ g/mL

The least number of significant digits used is three.

13.0 g/mL

```
13.25/1.02
        12.99019608
```

3. **Multiple Step Calculations**: - follow the multiply and divide rule.
   - Do NOT round off until the very LAST step.

Example 5: Calculate the final output energy in *Joules* if the equivalent mass of 5.3×10^{-3} kg is turned into energy along with an initial energy input of 4.15×10^{14} J. (Use $E = mc^2$ where $c = 3.00 \times 10^8$ m/s)

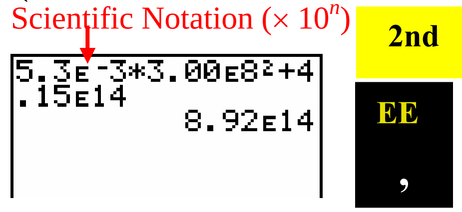
$E_{\text{output}} = mc^2 + E_{\text{input}}$

$E_{\text{output}} = (5.3 \times 10^{-3}\text{ kg})(3.00 \times 10^8\text{ m/s})^2 + 4.15 \times 10^{14}\text{ J}$

$E_{\text{output}} = 8.92 \times 10^{14}\text{ J}$

$\mathbf{E_{\text{output}} = 8.9 \times 10^{14}\text{ J}}$

Scientific Notation ($\times 10^n$)



Theoretical Result: - the supposed result of an experiment as calculated prior to the lab.

Experimental Result: - the actual measured result of an experiment.

Example 6: Determine the % Error and % Yield of an experiment if the theoretical result was 4.579 g and the experimental result was 4.272 g.

$$\%\text{ Error} = \frac{|\text{Theoretical} - \text{Experimental}|}{\text{Theoretical}} \times 100\% \qquad \%\text{ Yield} = \frac{\text{Experimental}}{\text{Theoretical}} \times 100\%$$

(100% is an exact number)

$$\%\text{ Error} = \frac{|4.579\text{ g} - 4.272\text{ g}|}{4.579\text{ g}} \times 100\% \qquad \%\text{ Yield} = \frac{4.272\text{ g}}{4.579\text{ g}} \times 100\%$$

% Error = 6.705% **% Yield = 93.30%**

Unit Conversions and Scientific Notations with Significant Digits:
- when changing units or converting scientific notation to standard notation or vice versa, ***the number of significant digits must remain the same as before***.

Example 7: Complete the following unit conversions.

a. 435. mL = **0.435** L **(left 3 places)**

b. 24. mg = **0.024** g **(left 3 places)**

c. 34500. mm = **345.00** dm **(left 2 places)**

d. 26.0 L = $\underbrace{26000}_{\text{too many significant; original measurement only has three digits.}}$ mL = **2.60×10^4** cm³ **(right 3 places)** (three significant digits)

e. 1854. dm = **0.01854** km **(left 5 places)**

f. 0.035 kg = $\underbrace{35000}_{\text{too many significant; original measurement only has two digits.}}$ mg = **3.5×10^4** mg **(right 6 places)** (two significant digits)

Dimensional Analysis: - commonly known as the ***unit factor method***.
- using units themselves to analyse their conversions or whether the right kind of procedure is used for calculations.
- unit factors have a bigger unit along with an equivalent smaller unit.
- final answer should keep the original number of significant digits.

Example 8: Convert 65.0 miles/h to km/h. (1 mile = 1.609344 km)

$$\frac{65.0\ \cancel{\text{miles}}}{1\ \text{hour}} \times \frac{1.609344\ \text{km}}{1\ \cancel{\text{mile}}} = 104.60736\ \text{km/h}$$ (round to 3 significant digits)

105. km/h or 105 km/h

Example 9: Convert 50. km/h to m/s.

$$\frac{50.\ \cancel{\text{km}}}{1\ \cancel{\text{hour}}} \times \frac{1000\ \text{m}}{1\ \cancel{\text{km}}} \times \frac{1\ \cancel{\text{hour}}}{3600\ \text{s}} = 13.8888889\ \text{m/s}$$ (round to 2 significant digits)

14. m/s or 14 m/s

Example 10: Convert 55. miles/gal to km/L. (1 gal = 3.785412 L)

$$\frac{55.\ \cancel{\text{miles}}}{1\ \cancel{\text{gal}}} \times \frac{1.609344\ \text{km}}{1\ \cancel{\text{mile}}} \times \frac{1\ \cancel{\text{gal}}}{3.785412\ \text{L}} = 23.38290257\ \text{km/L}$$ (round to 2 significant digits)

23. km/L or 23 km/L

Physical Change: - when the temperature of a substance is raised or lower, no new substance is formed. The only heat change is the **physical kinetic energy** of the substance as its molecules are moving or vibrating faster or slower.
- when a substance is changing phase, no new substance is formed. The only heat change is the **physical potential energy** of the substance as its intermolecular forces are being loosen or completely broken.

Heat Units: - the measuring units to measure heat or energy.

a. Joules: - the metric unit to measure heat or energy named after English physicist James Prescott Joule.

b. Calories: - the old imperial unit to measure heat or energy. (1 cal = 4.184 J)

Specific Heat Capacity: - the amount of heat needed to raise one gram of substance by one degree Celsius.
- the higher the specific heat capacity, the more the substance can "hold" the heat.
- units are in J/(g • °C) or kJ/(kg • °C)

Physical Kinetic Change

$$q = mc_P\Delta T$$

q = Change in Heat (J or kJ) **m = mass (g or kg)**
c_P = Specific Heat Capacity [J/(g • °C) or kJ/(kg • °C)]
ΔT = Change in Temperature (in °C)

Specific Heat Capacity of Some Common Substances (at 1.00 atm and 298.15 K)

Substance	Specific Heat Capacity J/(g • °C) or kJ/(kg • °C)	Substance	Specific Heat Capacity J/(g • °C) or kJ/(kg • °C)
Ice $H_2O_{(s)}$	2.01	Aluminum Al $_{(s)}$	0.90
Water $H_2O_{(l)}$	4.184	Carbon (graphite) C $_{(s)}$	0.709
Steam $H_2O_{(g)}$	2.0	Copper Cu $_{(s)}$	0.385
Ammonia $NH_{3\,(g)}$	2.06	Iron Fe $_{(s)}$	0.451
Methanol $CH_3OH_{(l)}$	2.53	Silver Ag $_{(s)}$	0.24
Ethanol $C_2H_5OH_{(l)}$	2.44	Gold Au $_{(s)}$	0.13

Example 11: How much energy in kJ, is needed to heat 100.0 g of water from 20.0°C to 80.0°C?

Since this question involves temperature (kinetic) change only, we need to use $q = mc_P\Delta T$.

c_P = 4.184 J/(g • °C)
m = 100.0 g H_2O
ΔT = 80.0°C – 20.0°C = 60.0°C
q = ?

$q = mC\Delta T$
$q = (100.0\ \cancel{g})(4.184\ J/(\cancel{g} \bullet \cancel{°C}))(60.0\cancel{°C}) = 25104$ J

q = 25.1 kJ

Example 12: A 20.0 g of an unknown element gave off 423 J of energy as it is cooled from 100°C to 10.0°C. Identify this unknown element by determining its specific heat capacity.

q = 423 J
m = 20.0 g
ΔT = 100.0°C – 10.0°C = 90.0°C
c_P = ?

$q = mc_P\Delta T$ $\quad \dfrac{q}{m\Delta T} = c_P$

$$c_P = \frac{(423\text{ J})}{(20.0\text{ g})(90.0^\circ\text{C})}$$

c_P = 0.235 J/(g • °C)

From the reference table above, the unknown substance is silver.

Example 13: Calculate the final temperature of a 10.0 g piece of copper wire at an initial temperature of 15.0°C when 3.27 kJ of heat is added.

q = 3.27 kJ
m = 10.0 g = 0.0100 kg of Cu
c_P = 0.385 kJ/(kg • °C)
T_i = 15.0°C

$\Delta T = T_f$ – 15.0°C = ?

$q = mc_P\Delta T$ $\quad \dfrac{q}{mc_P} = \Delta T$

$$\Delta T = \frac{(3.27\text{ kJ})}{(0.0100\text{ kg})(0.385\text{ kJ/(kg}\bullet^\circ\text{C}))}$$

ΔT = 849°C

$\Delta T = T_f$ – 15.0°C
849°C = T_f – 15.0°C

T_f = 849°C + 15.0°C

T_f = 864°C

Assignment

2.3 pg. 59 #1 to 3; pg. 61 #1 to 4; pg. 63 #1 to 11;
pg. 66−68 #8, 9, 13, 15, 18 to 42, 44, 45

