Unit 1: BASIC CHEMISTRY

Chapter 1: Scientific Method

1.3: The Scientific Method

Scientific Method: - a logical method to find solution 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 involving measurement with instruments.
4. **Theory**: - an idea that can explain a set of observation 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. Science can only tackle "testable" assumptions. Philosophical and Religious questions cannot be answer by science.

Chapter 2: Matter and Change

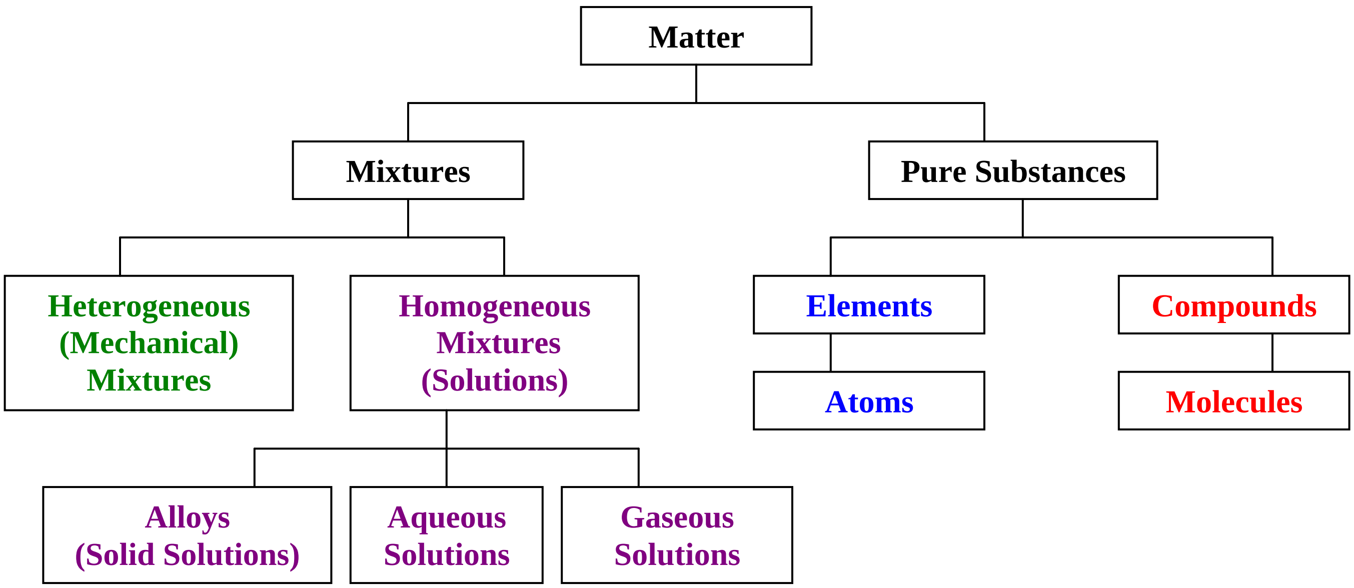
2.1 to 2.3: Matter, Mixtures & Elements and Compounds

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 mechanical mixtures which we can see its different components with the naked eye. An example of a heterogeneous mixture is a bag of assorted nuts. We can clearly see the different kind of nuts (walnuts, peanuts, chestnut, hazelnut ... etc.) in this bag. A **homogeneous (*homo* means the same) mixture** is 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. 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 atom. 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). 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.

The classification of matter is explained in a flow chart below.

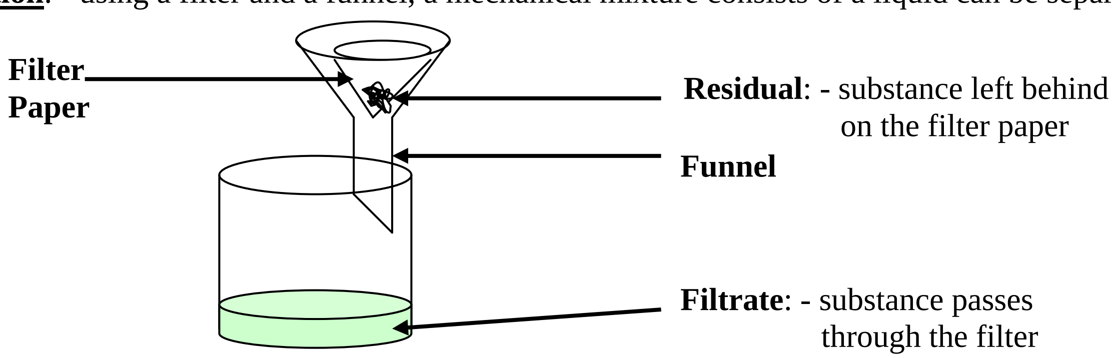


Matter: - anything that has a mass and occupies space.

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 a liquid 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	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.

2.4: Chemical Reactions

Chemical Reaction: - a process where **chemical change** has taken place.

Chemical Change: - a change where **New Substance(s)** are formed.

Five Evidences of a Chemical Change:

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.**
5. **Explosion!**

Reactants: - chemicals that goes into a reaction.

Products: - chemicals that are produced from a reaction.

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

Chemical Word Equation: - a chemical reaction written out in words.

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?

Iron	+ Oxygen	→ Rust
3.50 g	? g	5.1 g

Mass of Oxygen = 5.1 g – 3.50 g

Mass of Oxygen = 1.6 g

Assignment

1.3 pg. 17 #8 to 11
2.1 pg. 31 #1 to 4
2.2 pg. 34 #5, 6; pg. 35 #7 to 12
2.3 pg. 39 #13; pg. 40 #14 to 18
2.4 pg. 43 #19 to 23
Chapter 2 Review pg. 47–48 #24 to 43

Chapter 3: Scientific Measurement

3.1: The Importance of Measurement

Qualitative Measurements: - measurements that do not involve a numerical value.
(**Examples**: Colour, Odour, Heat Given off or Taken in, Type of Solid Formed … etc)

Quantitative Measurements: - measurements that do include a numerical value.
(**Examples**: Volume, Temperature, Mass, Time, … etc)

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

(1 to 9.999…) × 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)**

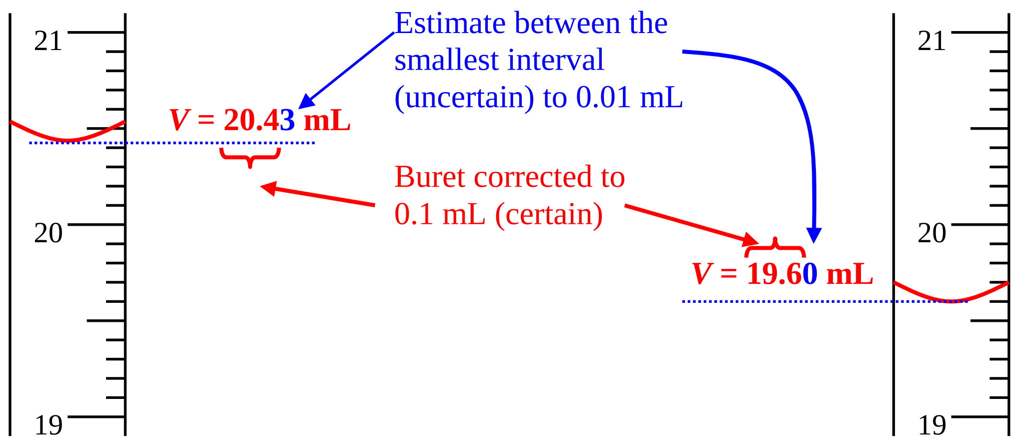
b. Mass of an Electron = 9.11×10^{-31} kg = **0.000 000 000 000 000 000 000 000 000 000 911 kg** **(moved 31 decimal places to the left)**

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)**

3.2: Uncertainty in Measurement

Uncertainty: - all measuring instruments have uncertainty sue to how the instrument was manufactured or reading error by the user.



Exact Number: - number that indicates no uncertainty. (Numbers in formulas; numbers written in words) or whole numbers with no decimal point of any objects – not measurements.)

Examples: **Two** chairs, $SA = \mathbf{4}\pi r^2$, **2500** atoms

Significant Digits: - digits used in the measurement plus one uncertain value.

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 contains a decimal point.

Example 1: State the number of significant digits for the following measurements.

a. 0.03 g — 1 significant digit
b. 0.030 g — 2 significant digits
c. 0.0304 g — 3 significant digits
d. 0.03040 g — 4 significant digits

e. 25 000 g — Exact Number
f. 9.300×10^4 m — 4 significant digits
g. 4.05×10^{-2} L — 3 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 2: 5.345 g + 0.42 g + 11.8 g

$$\begin{array}{r} 5.345 \\ 0.42 \\ +\,11.8 \\ \hline 17.565 \end{array}$$

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 3: $\dfrac{3.250\text{ g}}{1.4\text{ mL}}$

$\dfrac{3.250\text{ g}}{1.4\text{ mL}} = 2.321428571$ g/mL

The least number of significant digits used is two.

2.3 g/mL

```
3.250/1.4
          2.321428571
```

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

Example 4: 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)

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

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

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

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

Scientific Notation ($\times 10^n$)

```
5.3E-3*3.00E8²+4
.15E14
             8.92E14
```

2nd

EE

,

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%**

Assignment

3.1 pg. 53 #2 to 4

3.2 pg. 58 #5, 6; pg. 59 #7, 8; pg. 60 #9, 10; pg. 61 #11, 12; pg. 62 # 13 to 16

3.3: International System of Units

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)							

When converting units, the same number of significant digits must be preserved.

1 mL = 1 cm^3
× 1000 × 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. cm^3 = **0.026** L (26 cm^3 = 26 mL) **(left 3 places)**

e. 1854. cm = **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 × 10⁴** mg **(right 6 places)**

two significant digits

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 kg on Earth will have a mass of 50 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 1/6 of that on Earth.

3.4: Density

Density: - the amount of mass per unit of volume

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

Example 1: 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 Volume of the 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.)

```
4/3π5^3
        523.5987756
```

Substitute D and V to solve for m

To recall all digits of the previous answer

$$m = DV$$
$$m = (11.34\text{ g/cm}^3)(523.5987\ldots\text{ cm}^3)$$
$$m = 5.94 \times 10^3\text{ g or }5.94\text{ kg}$$

```
4/3π5^3
        523.5987756
11.34*Ans
        5937.610115
```

2nd

ANS

(–)

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

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

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

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

$$T_F = -34.6\,\text{F} \qquad T_K = 236.\text{ K}$$

Assignment

3.3 pg. 67 #17 to 22
3.4 pg. 71 #23, 24; pg. 72 #25, 26, 28
3.5 pg. 75 #30 to 35
Chapter 3 Review pg. 78–79 #36 to 61, 64 to 70

Chapter 5: Atomic Structure and the Periodic Table

5.1: Atoms

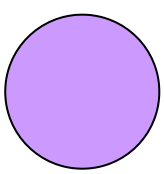
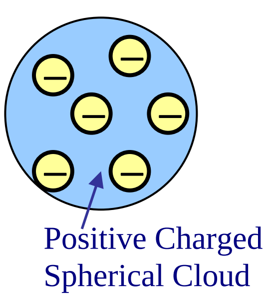
The practice of using symbols to represent elements can be traced back to the ancient Greek alchemists. Their purpose was to find a chemical recipe to make gold from other less valuable metals. (We now know that it is only possible now if we can change the number of protons in the nucleus).

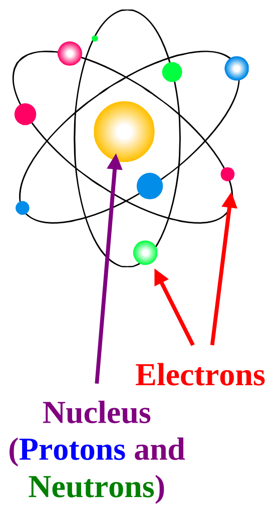
In 1808, a British scientist by the name of John Dalton published his theory of atoms that would have profound effects on the philosophy on chemistry and physics. (The word atom comes from the Greek word *atomos*, which means indivisible. A Greek philosopher Democritus in 5th-century BC first suggested this concept). The **Dalton's Atomic Theory** can be summarized as:

1. **All elements are made up of tiny particles called atoms.**
2. **The atoms of a particular element are identical. Different elements have different kind of atoms.**
3. **Atoms cannot be created or destroyed.**
4. **Chemical compounds are formed when different kinds of atoms combine together. A particular compound always has the same relative numbers and types of atoms.**
5. **Chemical reactions deal with the rearrangement of the atom, which changes the way they are combined together. There is no change to the atoms themselves in a chemical reaction.**

5.2: Structure of the Nuclear Atom

Since the time of Dalton's Atomic Theory, scientists had improved upon his model to better explain the structure of an atom. The following is a summary of the different atomic models.

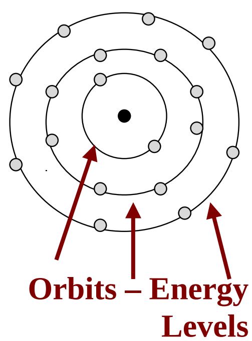
	**Dalton's Atomic Model:** In 1808, John Dalton proposed that all matter is made up of tiny particles called atoms. Atoms cannot be divided, created or destroyed. Different elements have different kinds of atoms. The difference is mainly due to the different atomic masses.
	**Plum Pudding Model:** In 1903, J.J. **Thomson** and Michael Faraday discovered **electrons** within an atom using a device called the **cathode ray tube**. Electrons are negatively charged subatomic particles with a **charge of –1**. The electrons were viewed as embedded in a positively charged spherical cloud. This is similar to the raisins distributed in a plum pudding. Robert **Millikan** used his **oil drop experiment** (by balancing the weight of an oil drop with electric force) to determine the elemental charge of the electron as 1.6×10^{-19} C and has a mass of 9.11×10^{-31} kg.



Nuclear Model:

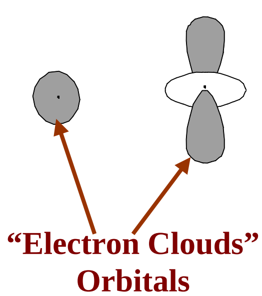
In 1912, Ernest **Rutherford** proposed the Nuclear Model for atoms after his famous **gold foil experiment**. Earlier to this time, **E. Goldstein** discovered the **positively charged (+1)** subatomic particles called **protons**. Rutherford proposed that the protons are packed tightly together at the centre of the atom called the **nucleus**. In 1932, James Chadwick discovered **neutrons (no charged)**. Together, they suggested that the **nucleus was made up of both protons and neutrons** (the bulk of the atomic mass) since electrons are very light compared to the masses of the protons and neutrons.

On the other hand, negatively charged electrons move around the nucleus because of their attraction with the positively charged nucleus (contains protons). Since the nucleus is very small, the circling electrons make up almost all of the volume of the atom. If the atom has a size of a football field, the nucleus is about the size of a small nail at the centre of the field.



The Bohr Model:

In 1913, Neil Bohr refined the Nuclear Model by suggesting that electrons move around the nucleus in specified **orbits**. These orbits are called **energy levels**. Electrons cannot exist between the orbits. The further the orbit is from the nucleus, the higher its energy level for the electrons in that orbit. This is very similar to the planetary model of our Solar system.



The Electron Cloud (Quantum Mechanics) Model:

This modern atomic model is very similar to the Bohr model. We still use the energy levels, however, the idea of orbits is modified into **orbitals**. **An orbital is a region of space where the electrons are most probably in**. Calculations of these orbital shapes involve advanced mathematics. Scientists use this model with the Molecular Orbital Theory to predict complex reactions and possible new chemical compounds.

Assignment
5.1 & 5.2 pg. 108 #1 to 3 & pg. 112 #4 to 6

5.3: Distinguishing Between Atoms

Subatomic Particles	Charge	Relative Mass	Actual Mass	Location
Electrons (e^-)	–1	1	9.11×10^{-31} kg	Region around the center of the atom
Protons (p^+)	+1	1836.12	1.67×10^{-27} kg	Centre of the atom called Nucleus
Neutrons (n)	0	1836.65	1.67×10^{-27} kg	Inside the Nucleus with the protons

Atomic Number and Atomic Mass:

We have looked at different atomic models. In high school chemistry, we deal mainly with the Bohr model. Recall from the Dalton's Atomic Theory, one of its points is that different elements have different atoms. The main difference between them is the **atomic mass**. This is the mass characteristic of a given element**. The atomic mass of an element is relative to the mass of the carbon atom (6 protons and 6 neutrons with an atomic mass of 12)**. It is usually located at the right, top corner or directly below each element on the Table of Elements. Atomic mass has a unit of amu (Atomic Mass Unit).

Because different elements have different atomic mass, the number of subatomic particles within an atom is also different for these elements. The **atomic number**, **a number assigned to each element based on its atomic mass**, is located at the top left corner of each element on the Table of Elements. **The atomic number is equated to the number of protons and electrons of that atom**. The **number of neutrons can be found by subtracting the atomic mass (rounded off whole number) with the atomic number**.

Atomic Number = Number of <u>Protons</u> and <u>Electrons</u> of an Atom

Number of Neutrons = Atomic Mass – Atomic Number

Example 1: State the Atomic Number, Atomic Mass, number of protons, neutrons, and electrons of the following elements.

17 35.45 **Cl** **Chlorine**	**Atomic Number = 17** (**17 p^+** and **17 e^-**) **Atomic Mass = 35.45** # of Neutron = 35.45 – 17 = **18 n**
1 1.01 **H** **Hydrogen**	**Atomic Number = 1** (**1 p^+** and **1 e^-**) **Atomic Mass = 1.01** # of Neutron = 1.01 – 1 = **0 n**

Note: Because any given atom has the same number of protons and electrons (same atomic number), **<u>all Atoms have a Net Charge of 0</u>**.

<u>**Isotopes**</u>:

Isotopes are atoms of an element with the same atomic number but a different mass because of a different number of neutrons. For a given mass of substance, there exist a certain percentage of isotopes. Some isotopes are stable. Others are unstable and they go through a decomposition process called **radioactive decay**.

Mass Number → A

Atomic Number → Z

$^{A}_{Z}X$ ← **Element Symbol**

A common example is the isotope $^{14}_{6}C$ (Carbon-14: Carbon with an atomic mass of 14 amu, which has 8 n, 6 p+ and 6 e^-). Naturally occur carbon contains 98.9 % of Carbon-12, 0.55% of Carbon-13 and 0.55% of Carbon-14. Chemists, physicists, archaeologists, geologists, and criminologists commonly use the carbon isotope. Because Carbon-14 is unstable and goes through radioactive decay at a definite rate, we can measure the amount of isotopes left in a substance to deduce its age. **Carbon-14 dating is a technique to date archaeological and geological findings by measuring the amount of Carbon-14 left in the artefacts**. Carbon-13 is used by chemists to assist in identifications of various chemical compounds.

Isotopes of other elements also have their uses in society. A tiny proportion of all water molecules (H_2O) compose of a hydrogen isotope called deuterium ($^{2}_{1}H$). Deuterium can be utilised as fuel in nuclear fusion reactors of the future. Other isotopes of various elements are used as **radiotracers**. These **are radioactive isotopes that can be introduced into organisms in food or drugs, and their pathways can be traced by monitoring their radioactivity**. These radiotracers have found their way into medical research. The list below shows some radiotracers and their medical applications.

Radiotracers	Area of the body examined	Radiotracers	Area of the body examined
$^{131}_{53}I$	Thyroid	$^{87}_{38}Sr$	Bones
$^{59}_{26}Fe$ and $^{51}_{24}Cr$	Red Blood Cells	$^{99}_{43}Tc$	Heart, Bones, Liver, and Lungs
$^{99}_{42}Mo$	Metabolism	$^{133}_{54}Xe$	Lungs
$^{32}_{15}P$	Eyes, Liver, Tumours	$^{24}_{11}Na$	Circulatory System

Average Atomic Mass Unit: - Average Mass of an atom and its isotopes which has taken account of their proportion of abundance (as stated on the Periodic Table of Elements).

Example 2: State the Average Atomic Mass Unit for hydrogen if it is made of 99.48% of $^{1}_{1}\text{H}$, 0.2400% of $^{2}_{1}\text{H}$, and 0.2800% of $^{3}_{1}\text{H}$.

$$\text{Average amu of Hydrogen} = \underbrace{(0.9948)(1)}_{\text{99.48\% of Atomic Mass 1}} + \underbrace{(0.002400)(2)}_{\text{0.2400\% of Atomic Mass 2}} + \underbrace{(0.002800)(3)}_{\text{0.2800\% of Atomic Mass 3}}$$

Average amu of Hydrogen = 1.008

5.4: The Periodic Table: Organizing the Elements

Electron Shells, Energy Levels and Valence Electrons:

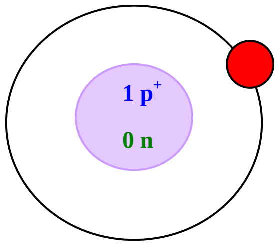
Recall from the Bohr Model studied earlier. It states that **electrons travel around the nucleus in specified orbits (electrons are quantum)**. These orbits are called **energy levels**. They can also be called **electron shells**. These orbits are very similar to the planets orbiting our sun. The only difference is that each orbit can accommodate more than one electron at a time. The following table shows the maximum number of electrons each successive "orbit" or energy level allows.

Energy Level	Maximum Number of Electrons Allowed
1st	2
2nd	8
3rd	8
4th	18
5th	18
6th	32
7th	32

To put electrons in the shells, we have to fill the first energy level until it is full before we can start filling the next energy level. If the second energy level is filled, then we can put electrons in the third energy level and so on. This process is repeated until all the electrons are used up. The following diagrams illustrate the point above.

Valence Electrons: - the electrons in the **outermost** shell.

Example 1: Draw the Bohr Energy Level diagram for the first 10 elements. State the number of protons, electrons, and neutrons as well as the location for each of these elements.

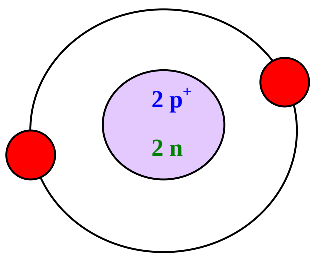


Element: Hydrogen (H)

Atomic Number: 1 **Atomic Mass: 1.01** **Nucleus: 1 p^+ and 0 n**

1st Energy Level: 1 e^- **(1 valence e^-)**
Total: 1 e^-

Location on the Period Table of Elements: First Row; Column IA

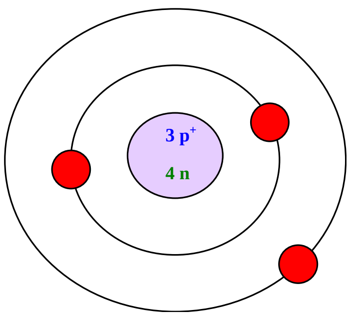


Element: Helium (He)

Atomic Number: 2 **Atomic Mass: 4.00** **Nucleus: 2 p^+ and 2 n**

1st Energy Level: 2 e^- **(2 valence e^- - Filled)**
Total: 2 e^-

Location on the Period Table of Elements: First Row; Column VIIIA

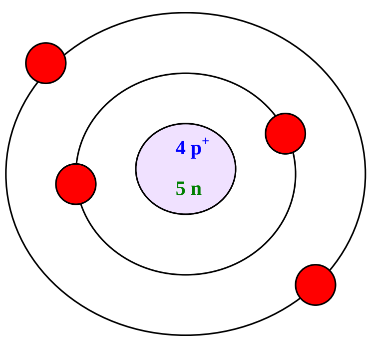


Element: Lithium (Li)

Atomic Number: 3 **Atomic Mass: 6.94** **Nucleus: 3 p^+ and 4 n**

2nd Energy Level: 1 e^- **(1 valence e^-)**
1st Energy Level: 2 e^-
Total: 3 e^-

Location on the Period Table of Elements: Second Row; Column IA

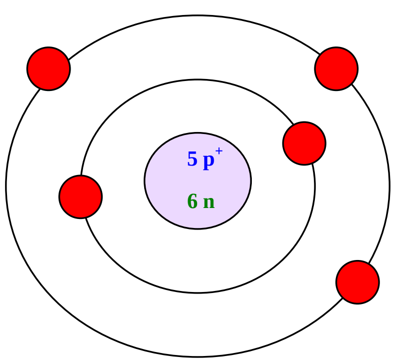


Element: Beryllium (Be)

Atomic Number: 4 **Atomic Mass: 9.01** **Nucleus: 4 p^+ and 5 n**

2nd Energy Level: 2 e^- **(2 valence e^-)**
1st Energy Level: 2 e^-
Total: 4 e^-

Location on the Period Table of Elements: Second Row; Column IIA

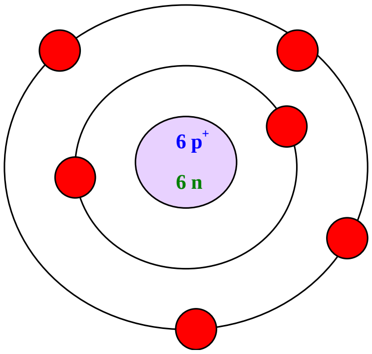


Element: Boron (B)

Atomic Number: 5 **Atomic Mass: 10.81** **Nucleus: 5 p^+ and 6 n**

2nd Energy Level: 3 e^- **(3 valence e^-)**
1st Energy Level: 2 e^-
Total: 5 e^-

Location on the Period Table of Elements: Second Row; Column IIIA

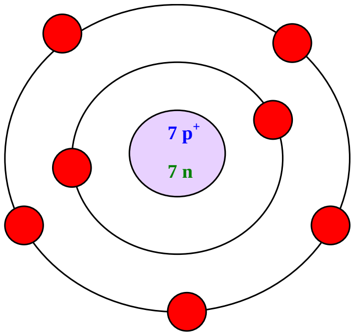


Element: Carbon (C)

Atomic Number: 6 Atomic Mass: 12.01 Nucleus: 6 p^+ and 6 n

2^{nd} Energy Level: 4 e^- (4 valence e^-)
1^{st} Energy Level: 2 e^-
Total: 6 e^-

Location on the Period Table of Elements: Second Row; Column IVA

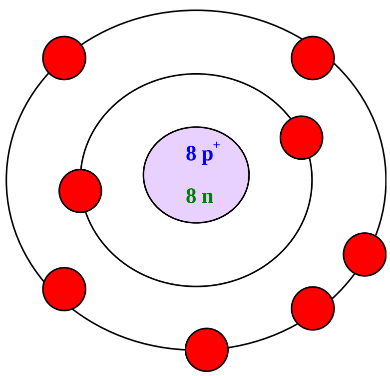


Element: Nitrogen (N)

Atomic Number: 7 Atomic Mass: 14.01 Nucleus: 7 p^+ and 7 n

2^{nd} Energy Level: 5 e^- (5 valence e^-)
1^{st} Energy Level: 2 e^-
Total: 7 e^-

Location on the Period Table of Elements: Second Row; Column VA

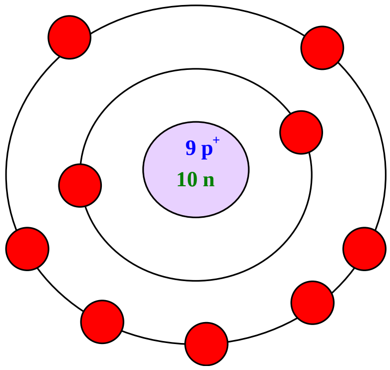


Element: Oxygen (O)

Atomic Number: 8 Atomic Mass: 16.00 Nucleus: 8 p^+ and 8 n

2^{nd} Energy Level: 6 e^- (6 valence e^-)
1^{st} Energy Level: 2 e^-
Total: 8 e^-

Location on the Period Table of Elements: Second Row; Column VIA

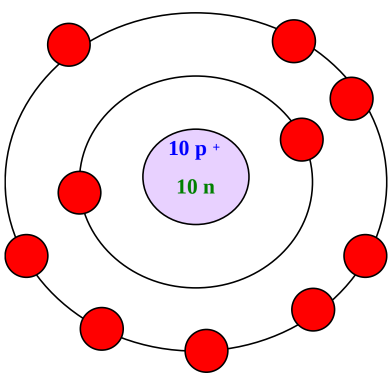


Element: Fluorine (F)

Atomic Number: 9 Atomic Mass: 19.00 Nucleus: 9 p^+ and 10 n

2^{nd} Energy Level: 7 e^- (7 valence e^-)
1^{st} Energy Level: 2 e^-
Total: 9 e^-

Location on the Period Table of Elements: Second Row; Column VIIA



Element: Neon (Ne)

Atomic Number: 10 Atomic Mass: 20.18 Nucleus: 10 p^+ and 10 n

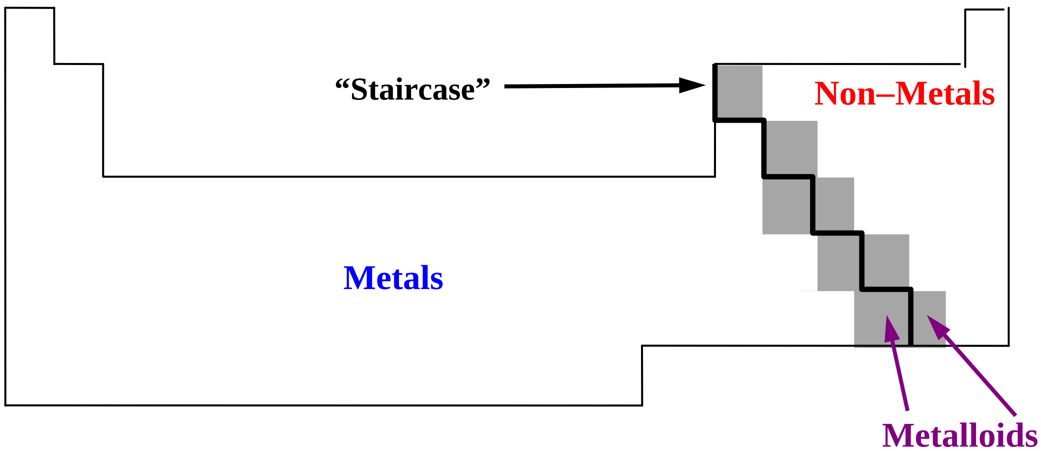
2^{nd} Energy Level: 8 e^- (8 valence e^- - Filled)
1^{st} Energy Level: 2 e^-
Total: 10 e^-

Location on the Period Table of Elements: Second Row; Column VIIIA

One way to remember the maximum number of electrons for each energy level is to look at the Periodic Table of Elements. There are 2 elements in the first row, hence 2 electrons are allowed in the first energy level. There are 8 elements each in the second and third rows, hence 8 electrons are allowed in each of the second and third energy level. This pattern repeats itself for higher energy levels.

Metals and Non-Metals

The 2 main categories of the Periodic Table of Elements are the **metals** and **non-metals**. They are divided by the "staircase" on the table. This "staircase" can be found at the element Boron extending down to the element Astatine. **Metals are the elements at the left side of the "staircase"**, and **non-metals are the elements at the right side of the "staircase"**.



Note: Hydrogen is placed at the metal side but is considered a non-metal.

Physical Properties: - are the properties or characteristics of a substance that can be change without involving the chemical change in its composition.

Physical Properties of Metals (with the exception of hydrogen):

1. Metals are mostly solids at room temperature (with the exception of mercury).
2. Metals are malleable (they can be hammered into thin sheets).
3. Metals are ductile (they can be pulled into wires).
4. Metals are good conducts of heat and electricity.
5. Metals are lustrous (shinny).

Physical Properties of Non-Metals:

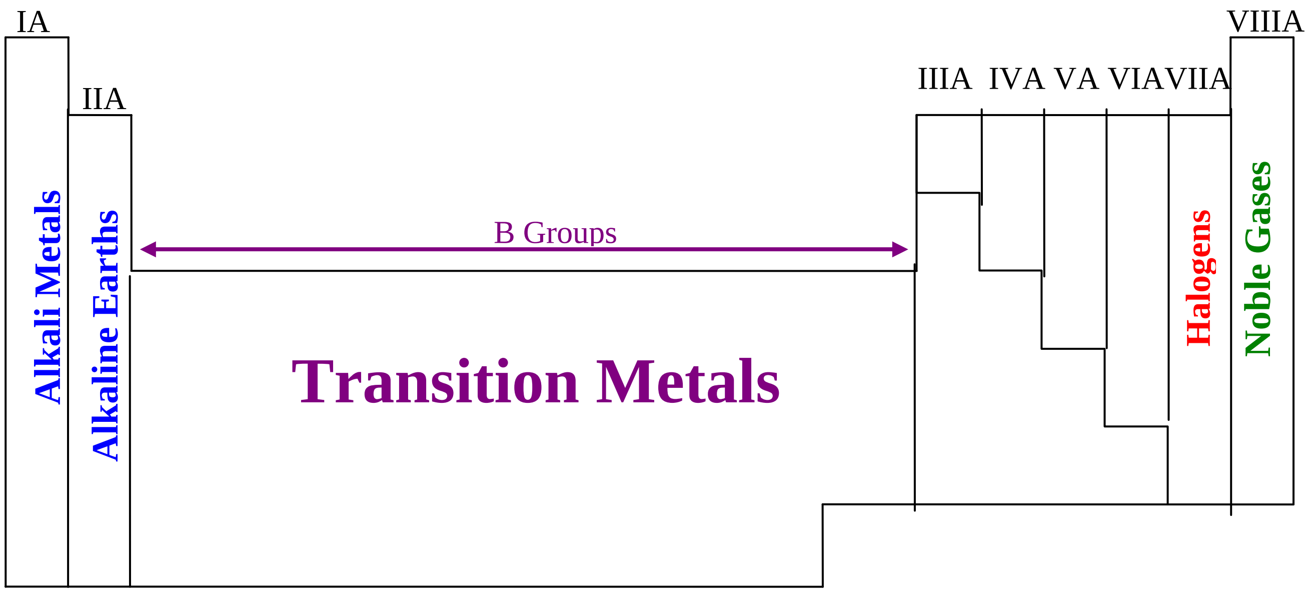
1. Non-metals are mostly gases or solids at room temperature (with the exception of bromide).
2. Non-metals usually do not have the physical properties of metals.

Some **elements near the "staircase" exhibit both the properties of metals and non-metals**. These elements are referred to as **metalloids**. An example is silicon. It is often used as a semiconductor material (an electrical conductor that can conduct and an insulate electricity). Other metalloids are boron, germanium, arsenic, antimony, tellurium, polonium, and astatine.

Periods and Groups: Chemical Properties of Elements

Chemical Properties: - the properties of a substance that involves a change in the organisation of atoms (mainly the sharing or transfer of electrons).

The shape of the Periodic Table of Elements is a structural way to organize elements. The vertical columns of the Table are called **groups** or **families**. As we have seen before, the column number is the same as the number of valence electrons of the elements. Since chemical properties depend greatly on the number of valence electrons, all elements within the same group or family must have similar chemical properties. We have already seen one such family, the noble gases. All elements of this group are non-reactive and very stable (recall the valence electron shell of these elements is full). The names of other families and their general chemical properties are listed below.



Groups or Families	Chemical Properties
Alkali Metals (IA)	very reactive metals
Alkaline Earth Metals (IIA)	less reactive than alkali metals
Halogens (VIIA)	very reactive non-metals
Noble Gases (VIIIA)	very stable; all are gaseous state at room temperature

Periods: - "rows" of elements that are identify by their highest energy level.
- the pattern of chemical properties "repeats" for every row.

Assignment

5.3 pg. 115 #7, 8; pg. 116 #9 to 11; pg. 117 #12, 13; pg. 120 #14, 15; pg. 121 #16 to 26
5.4 pg. 126 #28 to 30, 32
Chapter 5 Review pg. 129 #33 to 50, 52, 53