Unit 1: BASIC CHEMISTRY

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.

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

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)**

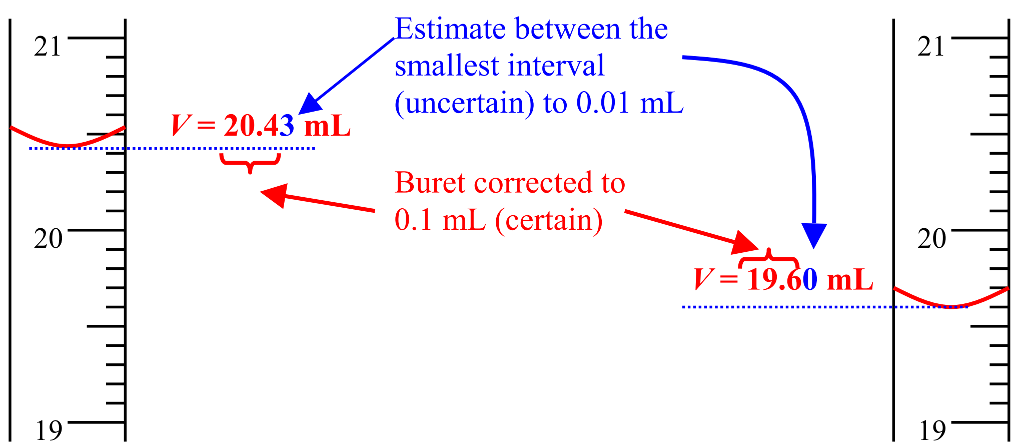
b. Mass of an Electron = 9.11×10^{-31} kg = **0.000 000 000 000 000 000 000 000 000 000 911 kg** **(moved 31 decimal places to the left)**

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)**

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)

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.

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 5 significant digits
f. 9.300×10^4 m 4 significant digits
g. 4.05×10^{-2} L 3 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 + 2.42 g + 1.8 g

```
  5.345
  2.42
 +1.8
 ------
  9.565
```

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

9.6 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
$\mathbf{E_{output} = 8.9 \times 10^{14}\text{ J}}$

Scientific Notation ($\times 10^n$)

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

2nd
EE
,

Theoretical Result: - the supposed result of an experiment according to pre-lab calculation.

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

Example 5: 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%**

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.

× 1000 **1 mL = 1 cm^3** × 1000

1 L = 1000 cm^3

Note: **1 m^3 ≠ 1 L**

$1\text{ m}^3 = 1\text{ m} \times 1\text{ m} \times 1\text{ m}$

$1\text{ m}^3 = 100\text{ cm} \times 100\text{ cm} \times 100\text{ cm}$

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 = $\underbrace{35000}$ mg = $\mathbf{3.5 \times 10^4}$ mg **(right 6 places)**

too many significant; original measurement only has two digits.

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.

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 pg. 67 #17 to 22

3.4: Density

Density: - the amount of mass per unit of volume

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

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 $\quad$ r = 5.00 cm $\quad$ 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

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

To recall all digits of the previous answer

```
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)

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

Example 2: With wind chill, Calgary can get down to −37°C. Convert the temperature to Fahrenheit and Kelvin.

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

$$T_F = \frac{9}{5}(-37) + 32 \qquad T_K = -37 + 273.15$$

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

$$\boldsymbol{T_F = -35 \text{ F}} \qquad \boldsymbol{T_K = 236 \text{ K}}$$

Assignment
3.4 pg. 71 #23, 24; pg. 72 #25, 26, 28
3.5 pg. 75 #30 to 35

Chapter 4: Problem Solving in Chemistry

4.2 & 4.3: Simple Conversion and More Complex Problems

Dimensional Analysis: - commonly known as unit factor method.
- using units to analyse unit conversion or whether the right kind of procedure is used for calculations.
- unit factors have 1 bigger unit along with equivalent smaller unit.
- should keep the original number of significant digits.

Example 1: 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

Example 2: 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

Example 3: 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

Assignment

4.2 pg. 93 #9, 10; pg. 94 #11 to 13; pg. 95 #15 to 19
4.3 pg. 100 #28 to 31

Assignment

Chapter 3 Review pg. 78–79 #36 to 61, 64 to 70
Chapter 4 Review pg. 103–104 #33 to 50, 52 to 55

Chapter 2: Matter and Change

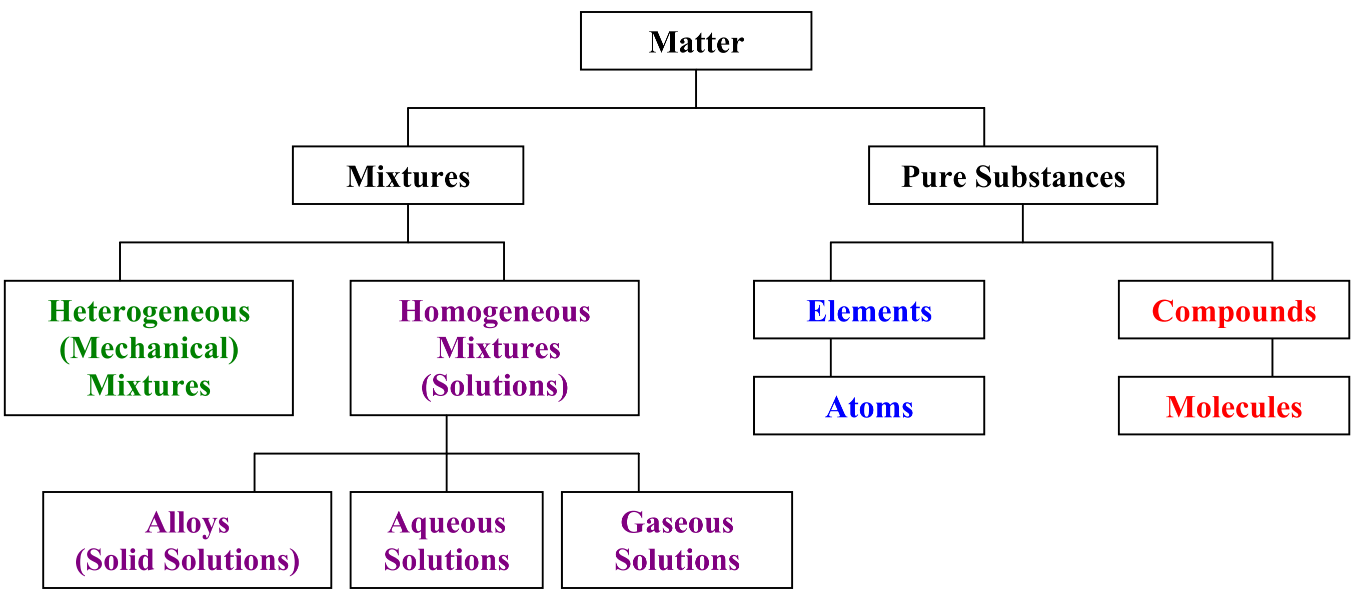
2.1 to 2.3: Matter, Mixtures & Elements and Compounds

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 2 hydrogen atoms and an oxygen atom.

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

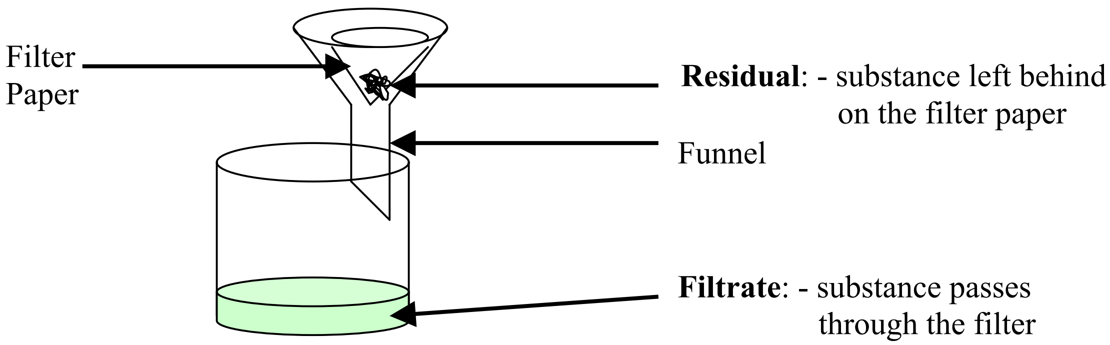


<u>Matter</u>: - anything that has a mass and occupies space.

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 being dissolve. **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 (a particle that is made up of more than one kind of 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	e) carbon dioxide	i) salt	m) propane
b) water	f) chlorine	j) nickel	n) baking soda
c) ammonia	g) ethanol	k) gold	o) uranium
d) oxygen	h) charcoal (carbon)	l) neon	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

   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.

Assignment

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: 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
2.4 pg. 43 #19 to 23

Chapter 5: Atomic Structure and the Periodic Table

5.1: Atoms

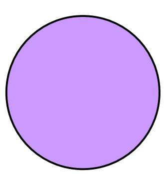
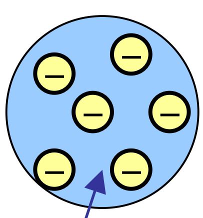
The practice of using symbols to represent elements can be traced back to the ancient Greek alchemists. Their purpose was to find a chemical recipe to make gold from other less valuable metals. (We now know that it is only possible now if we can change the number of protons in the nucleus).

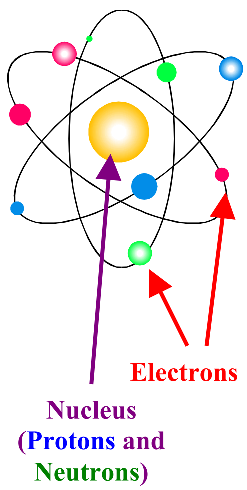
In 1808, a British scientist by the name of John Dalton published his theory of atoms that would have profound effects on the philosophy on chemistry and physics. (The word atom comes from the Greek word *atomos*, which means indivisible. A Greek philosopher Democritus in 5th-century BC first suggested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

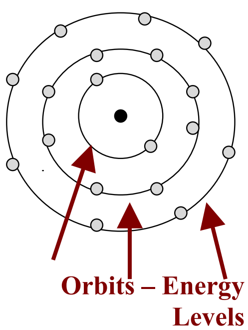
	**Dalton's Atomic Model:** In 1808, John Dalton proposed that all matter is made up of tiny particles called atoms. Atoms cannot be divided, created or destroyed. Different elements have different kinds of atoms. The difference is mainly due to the different atomic masses.
	**Plum Pudding Model:** In 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:

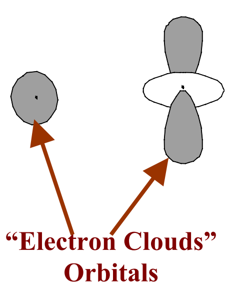
In 1912, Ernest **Rutherford** proposed the Nuclear Model for atoms after his famous **gold foil experiment**. Earlier to this time, E. Goldstein discovered the **positively charged (+1)** subatomic particles called **protons**. Rutherford proposed that the protons are packed tightly together at the 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.

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 Protons and Electrons 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), **all Atoms have a Net Charge of 0**.

Isotopes:

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 $\rightarrow$ A

Atomic Number $\rightarrow$ Z

$${}^{A}_{Z}X \leftarrow \textbf{Element Symbol}$$

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

Average Atomic Mass Unit: - Average Mass of an atom and its isotopes 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 $^{1}_{1}\text{H}$, 0.24% of $^{2}_{1}\text{H}$, and 0.28% of $^{3}_{1}\text{H}$.

$$\text{Average amu of Hydrogen} = \underbrace{(0.9948)(1)}_{\text{99.48\% of Atomic Mass 1}} + \underbrace{(0.0024)(2)}_{\text{0.24\% of Atomic Mass 2}} + \underbrace{(0.0028)(3)}_{\text{0.28\% of Atomic Mass 3}}$$

Average amu of Hydrogen = 1.008

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: The Periodic Table: Organizing the Elements

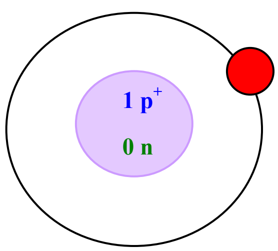
Electron Shells, Energy Levels and Valence Electrons:

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.

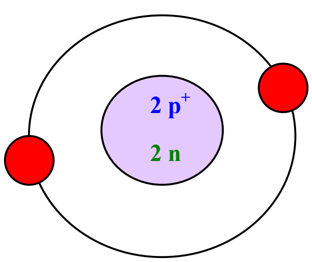


Element: Hydrogen (H)

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

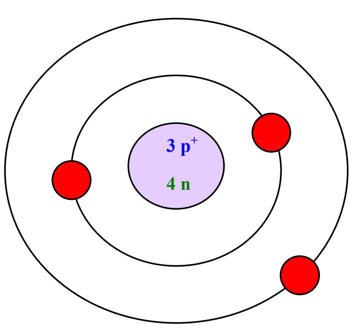


Element: Helium (He)

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

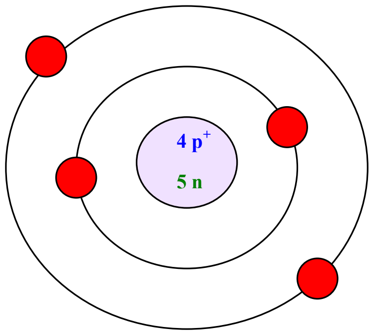


Element: Lithium (Li)

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

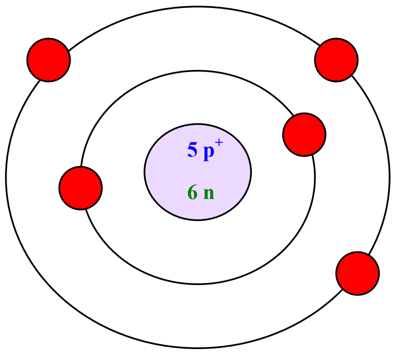


Element: Beryllium (Be)

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

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

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

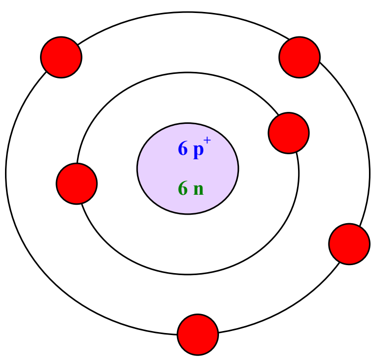


Element: Boron (B)

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

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

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

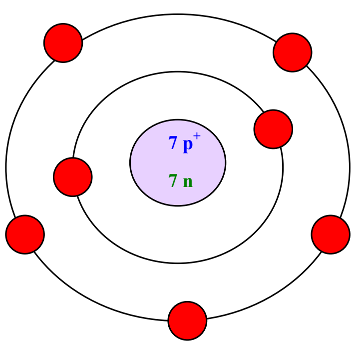


Element: Carbon (C)

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

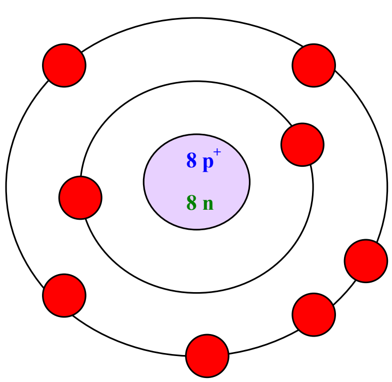


Element: Nitrogen (N)

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

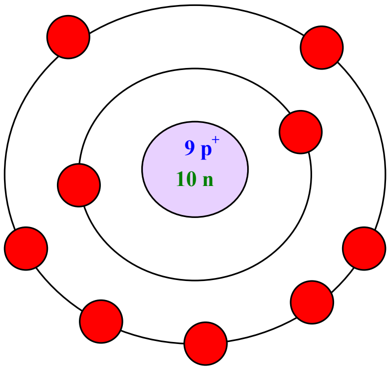


Element: Oxygen (O)

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

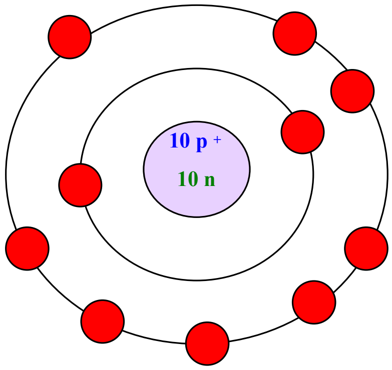


Element: Fluorine (F)

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

2nd Energy Level: 7 e^- (7 valence e^-)
1st 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**

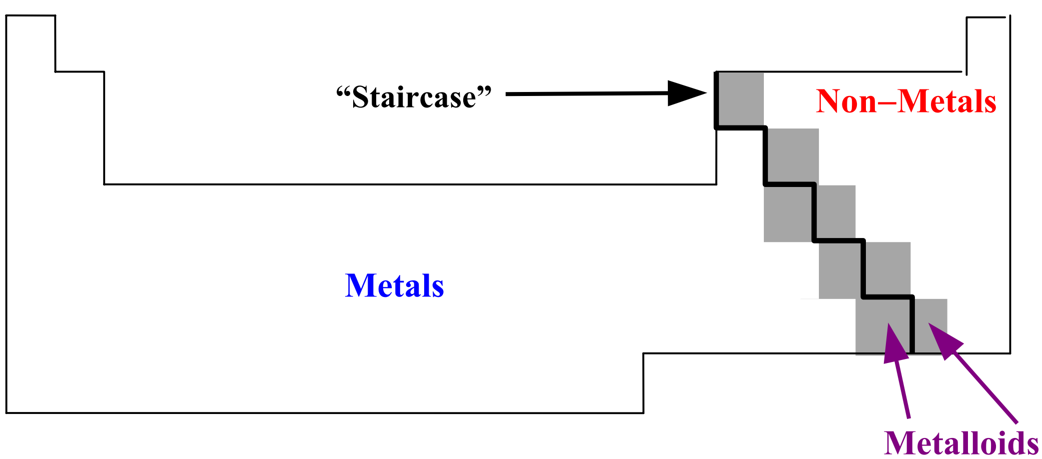
2nd Energy Level: 8 e^- (8 valence e^- - Filled)
1st Energy Level: 2 e^-
Total: 10 e^-

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:

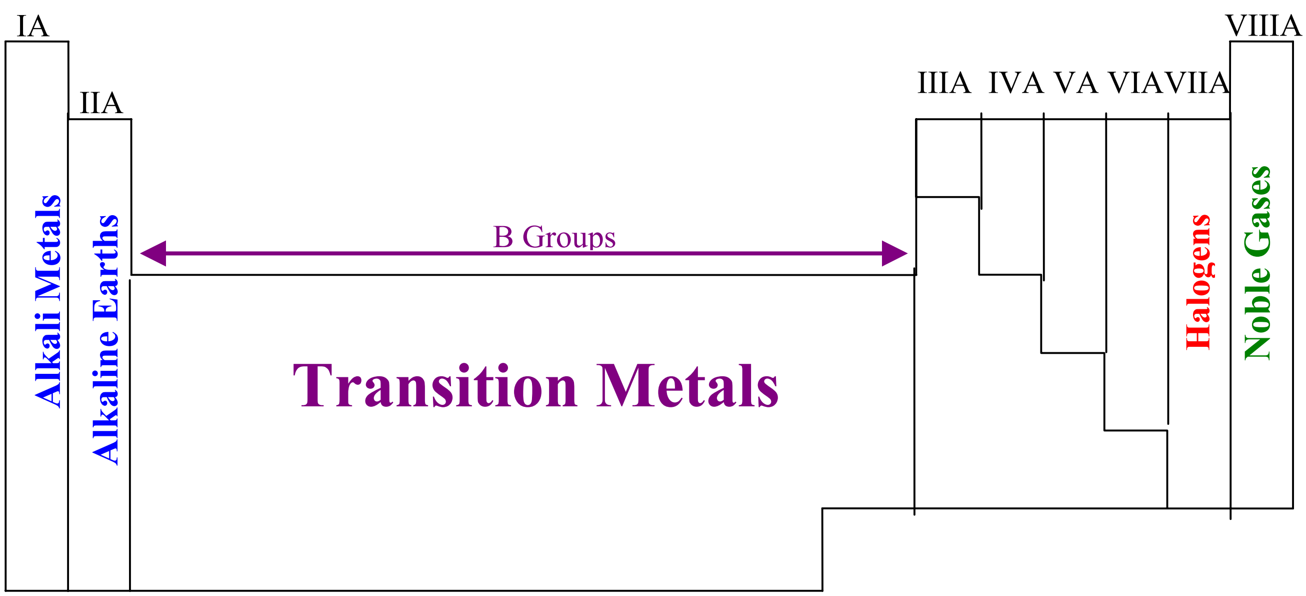
1. Non-metals are mostly gases or solids at room temperature (with the exception of bromide).
2. Non-metals usually do not have the physical properties of metals.

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 organise 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
5.4 pg. 126 #28 to 30, 32

Assignment
Chapter 2 Review pg. 47–48 #24 to 43
Chapter 5 Review pg. 129 #33 to 50, 52, 53

Chapter 6: Chemical Names and Formulas

6.1: Introduction to Chemical Bonding

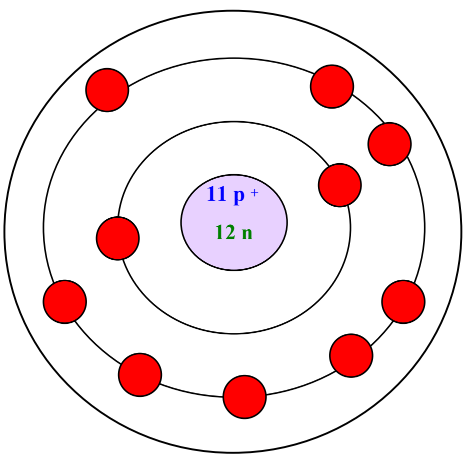
Molecules: - basic unit of a compound.
- contain at least two atoms of the same or different kind of elements.

Ions: - when atoms lose or gain electrons, they attain a positive or negative charge.

1. **Cations**: - **positive** charged ions (atoms that lose electrons).
   - naming cation (element name follow by "ion")

Example 1: Draw the energy level diagrams for the following cations.

a. Sodium ion = Na^+ (11 p^+ and 10 e^-)



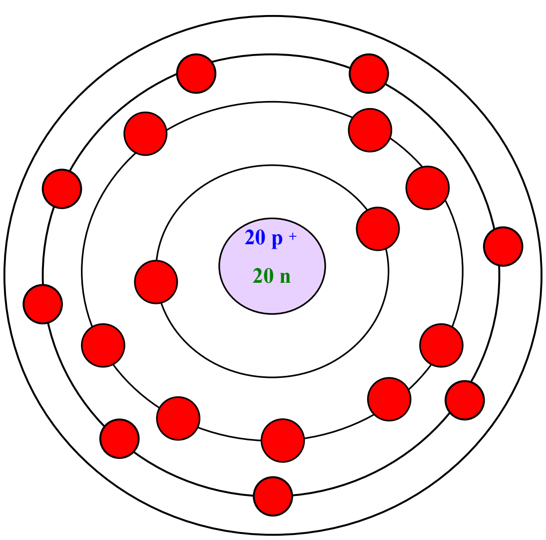
Sodium Ion (Na^+)

Atomic Number: 11 **Atomic Mass: 22.99** **Nucleus: 11 p^+ and 12 n**

2^{nd} Energy Level: 8 e^- (8 valence e^- - Filled)
1^{st} Energy Level: 2 e^-
Total: 10 e^- **Net Charge = 1+**

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

b. Calcium ion = Ca^{2+} (20 p^+ and 18 e^-)



Calcium Ion (Ca^{2+})

Atomic Number: 20 **Atomic Mass: 40.08** **Nucleus: 20 p^+ and 20 n**

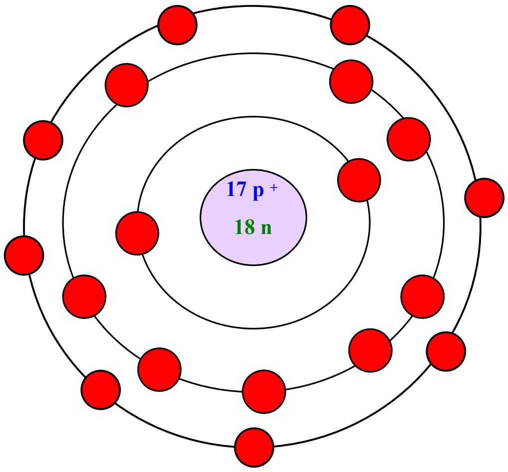
3^{rd} Energy Level: 8 e^- (8 valence e^- - Filled)
2^{nd} Energy Level: 8 e^-
1^{st} Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 2+**

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

2. **Anions**: - **negative** charged ions (atoms that gain electrons).
- naming anion (first part of element name follow by suffix ~*ide*)

Example 2: Draw the energy level diagrams for the following anions.

a. Chloride = Cl^- (17 p^+ and 18 e^-)



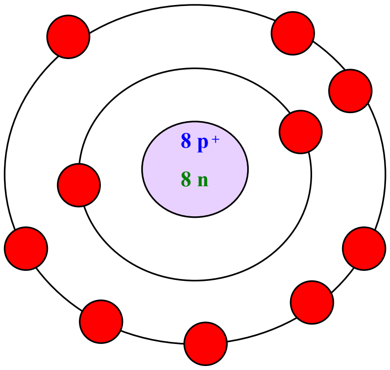
Chloride (Cl^-)

Atomic Number: 17 **Atomic Mass: 35.45** **Nucleus: 17 p^+ and 18 n**

3^{rd} Energy Level: 8 e^- (8 valence e^- - Filled)
2^{nd} Energy Level: 8 e^-
1^{st} Energy Level: 2 e^-
Total: 18 e^- **Net Charge = 1–**

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

b. Oxide = O^{2-} (8 p^+ and 10 e^-)



Oxide (O^{2-})

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

2^{nd} Energy Level: 8 e^- (8 valence e^- - Filled)
1^{st} Energy Level: 2 e^-
Total: 10 e^- **Net Charge = 2–**

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

Octet Rule: - the tendency for electrons to fill the second and third energy levels (8 valence e^-) to achieve stability.

6.2: Representing Chemical Compounds

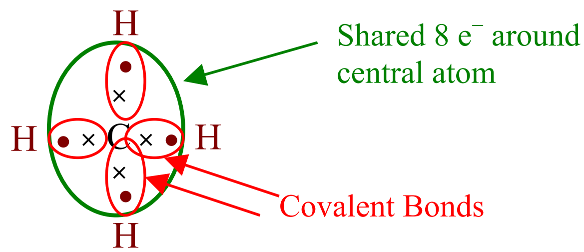
Chemical Formula: - displays the different atoms and their numbers in the smallest representative unit of a substance (includes all molecular compound, monoatomic, binary and polyatomic elements).

Molecular Formula: - chemical formula that represents the actual molecule of a molecular compound.

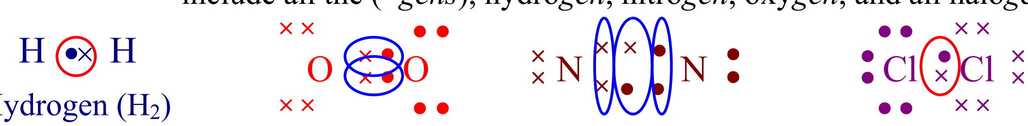
Formula Units: - chemical formulas that represent the basic unit of an ionic compound (compounds that made up of a metal and a non-metal)

Molecular Compound: - when a non-metal combines with a non-metal.
- forms **covalent bonds** (electrons are "share" between atoms).
- forms **non-electrolytes** (do not dissociate into ions when dissolve in water)

Example: CH_4



Binary Elements: - non-metals that come in pairs of atoms as molecules to achieve stability.
- include all the (~*gens*), hydro*gen*, nitro*gen*, oxy*gen*, and all halo*gen*s.



Hydrogen (H_2) Oxygen (O_2) Nitrogen (N_2) Chlorine (Cl_2)

Polyatomic Elements: - non-metals that comes in groups of 4 or 8 to achieve stability.
- **P_4 (Phosphorus)** and **S_8 (Sulphur)**

Diatomic Elements **(all the ~gens, including Halogens -** second last column of the Periodic Table**)**	**Polyatomic Elements**	**Monoatomic Elements**
Hydrogen (H_2), Oxygen (O_2), Nitrogen (N_2), Fluorine (F_2), Chlorine (Cl_2), Bromine (Br_2), Iodine (I_2), Astatine (At_2)	Phosphorus (P_4) Sulphur (S_8)	All other Elements. **Examples**: Helium (He), Iron (Fe), Calcium (Ca), Silver (Ag), Mercury (Hg)

Note: **Students should memorize all the diatomic and polyatomic elements. They are the only exceptions. All other elements are monoatomic.** A lot of the symbols are recognisable from the name of the elements (Zinc: Zn; Carbon: C; Aluminium: Al). Some of them look somewhat different. This is because the symbols came from the elements' Latin names (Silver: Ag for "Argentum"; Gold:

Au for "Aurum"). To save time, students should also familiarise themselves with the whereabouts of the elements on the Table.

Law of Definite Proportion: - the same compound always contains exactly the same proportion of elements by mass. (*Proust*)

Example 1: Water contains about 8 parts oxygen to 1 part hydrogen by mass. A 192 g of unknown liquid composed of hydrogen and oxygen contains 12 g of hydrogen. Is the unknown liquid water? Justify your response.

192 g total − 12 g of hydrogen = 180 g of oxygen

$$\frac{180\text{ g oxygen}}{12\text{ g hydrogen}} = \frac{15\text{ parts oxygen}}{1\text{ part hydrogen}}$$

Since the ratio between oxygen and hydrogen is 8:1 in water, the unknown liquid is NOT water.

Law of Multiple Proportion: - when two elements form a series of compounds, the ratios of the masses of the second element that combine with the first element can always be reduced to small whole numbers. (*Dalton*)

Example 2: State the ratios of hydrogen between the following hydrocarbon compounds.

Hydrocarbons	Mass of Hydrogen per 1 g of Carbon
Compound A	2.973 g
Compound B	3.963 g
Compound C	4.459 g

$$\frac{A}{B} = \frac{2.973\text{ g}}{3.963\text{ g}} = 0.7501892506 \approx 0.75 \qquad \frac{A}{B} = \frac{3}{4}$$

$$\frac{B}{C} = \frac{3.963\text{ g}}{4.459\text{ g}} = 0.8887642969 \approx 0.888\ldots \qquad \frac{B}{C} = \frac{8}{9}$$

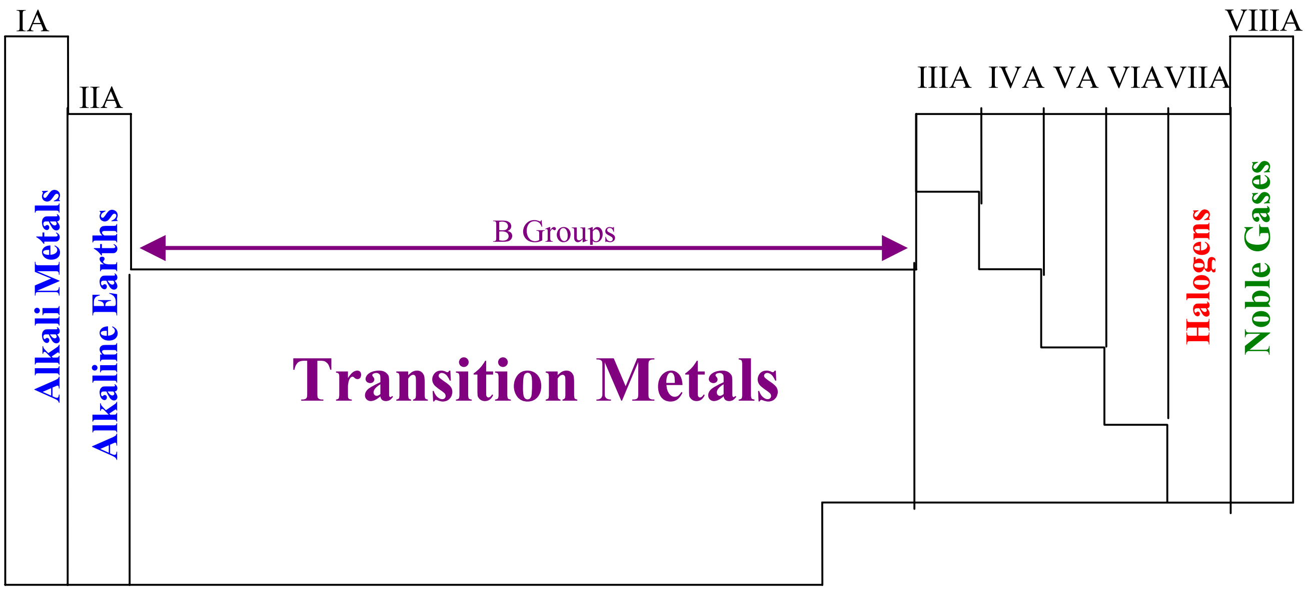
$$\frac{C}{A} = \frac{4.459\text{ g}}{2.973\text{ g}} = 1.49983182 \approx 1.5 \qquad \frac{C}{A} = \frac{3}{2}$$

Assignment

6.1 pg. 137 #1 to 9

6.2 pg. 142 #10, 11, 14 and 15

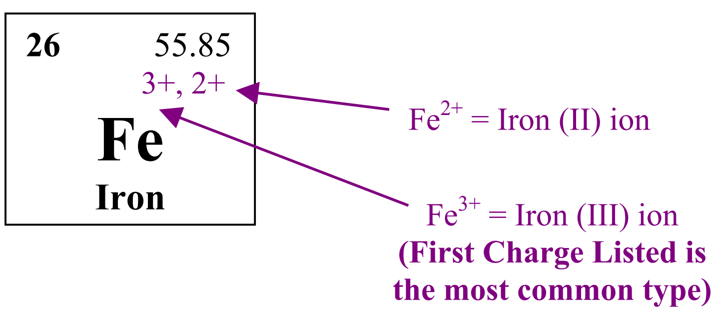
6.3: Ionic Charges



Groups or Families	Chemical Properties
Alkali Metals (IA)	very reactive; forms ions with +1 charge when react with non-metals
Alkaline Earth Metals (IIA)	less reactive than alkali metals; forms ions with +2 charge when react with non-metals
Halogens (VIIA)	very reactive; form ions with –1 charge when react with metals; all form diatomic molecules
Noble Gases (VIIIA)	very stable; do not form ions; monoatomic gas at room temperature

Transition Metals (1B to 10B): - groups and periods of metals that can have **varying charges**.
- use **Roman Numerals** as part of their ionic names.

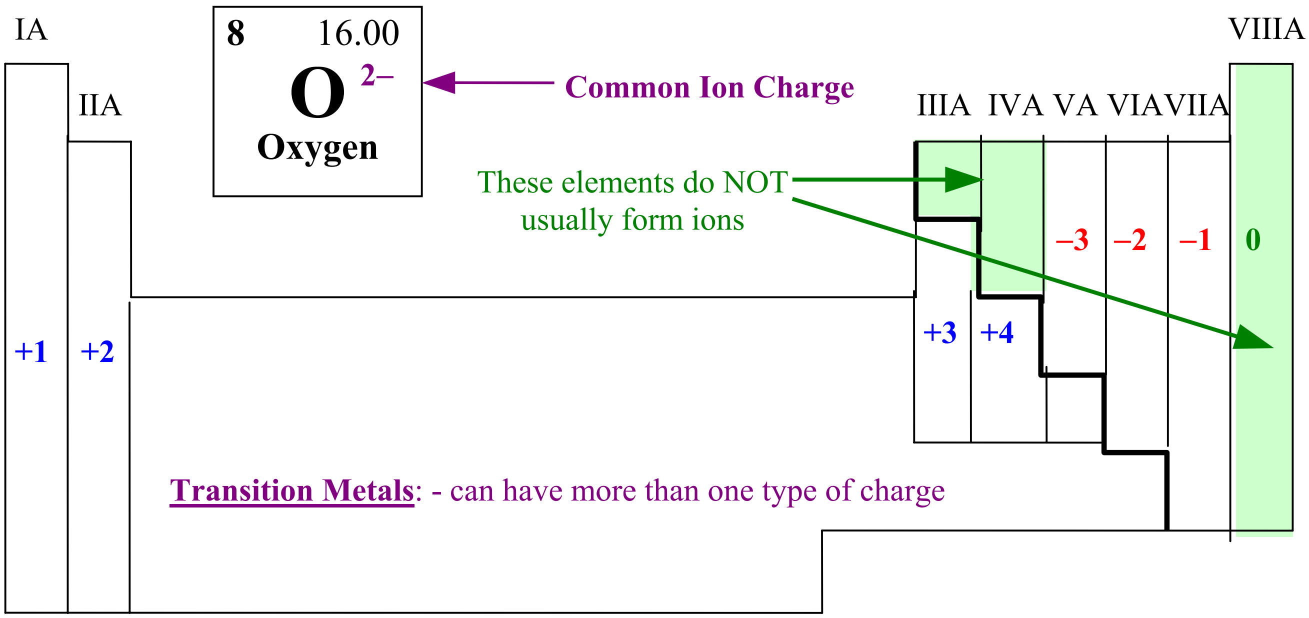
Example: Fe^{3+} and Fe^{2+} ions



Chemical Properties of Metals and Non-Metals:

1. **Metals lose electrons to become positive ions – cations**.
2. **Non-Metals gain electrons to become negative ions – anions**.
3. Hydrogen usually loses an electron to become a H^+ ion. However, it can sometimes gain an electron to become H^- (Hydride).
4. **The last column of the Table of Elements does not usually form ions**. These elements are called the **Noble Gases** (Helium, Neon, Argon, Krypton, Xenon, and Radon).

The number of electrons an atom loses or gains depends on which column (vertical) the element is at the Table.



The reason that noble gases (column VIIIA) do not form ions is because their outermost shells are filled with the maximum number of electrons allowed. That is why we call this group of elements "noble gases". They do not form ions because they are stable. Hence we use the word "noble" to describe them. All the other elements form ions because they want to achieve stability like the noble gases. If you observe carefully, oxide has the same number of electrons as the nearest noble gas, neon. On the other hand, calcium ion has the same number of electrons as the nearest noble gas, argon. In terms of stability, which is another word for lower energy state, these ions are more stable than their respective atoms.

Since the number of valence electrons of an atom is the same as its column number, all the elements of column IA have 1 valence electron. As we see with lithium, all they have to do is to lose that valence electron to achieve a noble gas "like" state. For elements in column IIA, they all have 2 valence electrons. Hence, they lose 2 electrons to acquire stability and become ions with a net charge of +2. The following table summarises these points.

Column	Number of Valence Electrons	Methods to achieve a Stable State	Net Charge of Ions
IA	1	lose 1 electron ~~or gain 7 electrons~~	+1
IIA	2	lose 2 electrons ~~or gain 6 electrons~~	+2
IIIA	3	lose 3 electrons ~~or gain 5 electrons~~	+3
IVA	4	lose 4 electrons ~~or gain 4 electrons~~	+4
VA	5	~~lose 5 electrons or~~ gain 3 electrons	–3
VIA	6	~~lose 6 electrons or~~ gain 2 electrons	–2
VIIA	7	~~lose 7 electrons or~~ gain 1 electron	–1
VIIIA	8	already has the maximum number of electrons allowed in the outermost electron shell.	0

Monoatomic Ions: - ions that came from a single atom (include metal cations and non-metal anions).
- monoatomic anion ends with suffix ~ide.

Examples: Na^+ = sodium ion, Cl^- = chloride, Pb^{4+} = lead (IV) ion, Zn^{2+} = zinc ion

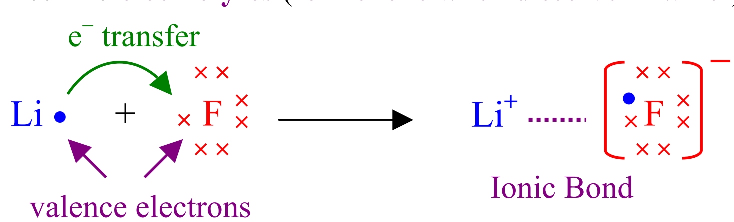
Polyatomic (Complex) Ions: - ions that contain many atoms.
- mostly anions (except NH_4^+ = ammonium ion).
- most ends with suffixes ~*ate* or ~*ite* (some ends with suffix ~*ide*).

Examples: CO_3^{2-} = carbonate, $Cr_2O_7^{2-}$ = dichromate, OH^- = hydroxide, SO_3^{2-} = sulfite

6.4: Ionic Compound

Ionic Compound: - when a metal combines with a non-metal.
- forms **ionic bonds** (electrons are "stolen" or "transfer" from one atom to another).
- dissociates into **electrolytes** (forms ions when dissolve in water)

Example: LiF



Nomenclature: - a naming system

IUPAC: - International Union of Pure and Applied Chemistry.
- an organisation that oversees the standard nomenclature of all chemicals.

Nomenclature of Ionic Compounds

1. **Balance the Cation and Anion Charges.**
2. Use **brackets** for **multiple Complex Ions (Polyatomic Ions)**.
3. When naming, use ***~ide*** for the **non-metal anions**.
4. **Metals** that can have **two or more different charges** must use **Roman Numerals** in the names.

Example 1: Write the chemical formulas for the following.

a. sodium chloride

Na^+ and Cl^- ⇒ **NaCl**

Need 1 Na^+ & 1 Cl^- to balance charges

b. calcium fluoride

Ca^{2+} and F^- ⇒ **CaF_2**

Need 1 Ca^{2+} & 2 F^- to balance charges

c. ammonium sulfate

NH_4^+ and SO_4^{2-} ⇒ **$(NH_4)_2SO_4$**

2 (NH_4^+) & 1 SO_4^{2-} to balance charges

d. magnesium hydroxide

Mg^{2+} and OH^- ⇒ **$Mg(OH)_2$**

Need 1 Mg^{2+} & 2 (OH^-) to balance charges

e. tin (IV) sulfite

Sn^{4+} and SO_3^{2-} ⇒ **$Sn(SO_3)_2$**

Need 1 Sn^{4+} (IV means 4+ charge) & 2 (SO_3^{2-}) to balance charges

f. aluminium oxide

Al^{3+} and O^{2-} ⇒ **Al_2O_3**

Need 2 Al^{3+} & 3 O^{2-} to balance charges

Oxyanions: - a series of polyatomic ions that contains different number of oxygen atoms.

SO_4^{2-}	sulf***ate***	ClO_4^-	***per***chlor***ate***
SO_3^{2-}	sulf***ite***	ClO_3^-	chlor***ate***
NO_3^-	nitr***ate***	ClO_2^-	chlor***ite***
NO_2^-	nitr***ite***	ClO^-	***hypo***chlor***ite***

Example 2: Name the following oxyanions.

a. BrO_4^-
BrO_3^-
BrO_2^-
BrO^-

BrO_4^-	***per***brom***ate***
BrO_3^-	brom***ate***
BrO_2^-	brom***ite***
BrO^-	***hypo***brom***ite***

b. IO_4^-
IO_3^-
IO_2^-
IO^-

IO_4^-	***per***iod***ate***
IO_3^-	iod***ate***
IO_2^-	iod***ite***
IO^-	***hypo***iod***ite***

Example 3: Name the following ionic compounds.

a. Na_2S

Na^+ = sodium
S^{2-} = sulfide
sodium sulfide

b. $KMnO_4$

K^+ = potassium
MnO_4^- = permanganate
potassium permanganate

c. $Li_2S_2O_3$

Li^+ = lithium
$S_2O_3^{2-}$ = thiosulfate
lithium thiosulfate

d. $CuCr_2O_7$

Cu^{2+} = copper (II)
$Cr_2O_7^{2-}$ = dichromate
copper (II) dichromate

e. $Fe_3(BO_3)_2$

Fe^{2+} = iron (II)
BO_3^{3-} = borate
iron (II) borate

f. $Ti(BrO_3)_3$

Ti^{3+} = titanium (III)
BrO_3^- = bromate
titanium (III) bromate

Hydrate: - ionic compounds sometimes come with water molecule locked in their crystal form.
- naming contains the ionic compound name with a prefix follow by the word "*hydrate*".

Prefixes for Hydrates

1 - mono	4 - tetra	7 - hepta	10 - deca
2 - di	5 - penta	8 - octa	
3 - tri	6 - hexa	9 - nona	

Example: $CuSO_4 \bullet 5H_2O$ copper (II) sulfate pentahydrate

Assignment

6.3 pg. 145 #16, 17; pg. 146 #18, 19; pg. 148 #20, 22 and 23
6.4 pg. 151 #24, 25; pg. 153 #26, 27; pg. 155 #28, 29; pg. 156 #30 to 36

6.5: Molecular Compounds and Acids

Nomenclature of Molecular Compounds

1. Do NOT use charges to balance subscripts. Use prefixes to name or write the formula's subscripts.
2. If the first element has one atom in the molecule, do NOT use *mono~* as a prefix.
3. The last element uses the suffix *~ide*.

Prefixes for Binary Molecular Compounds

1 - mono	4 - tetra	7 - hepta	10 - deca
2 - di	5 - penta	8 - octa	
3 - tri	6 - hexa	9 - nona	

Example 1: Name the following molecular compounds.

a. CO
1 Carbon and 1 Oxygen
Carbon monoxide

b. CO_2
1 Carbon and 2 Oxygen
Carbon dioxide

c. N_2O_4
2 Nitrogen and 4 Oxygen
dinitrogen tetraoxide

Example 2: Provide the chemical formula for the following compounds

a. sulfur trioxide
1 S and 3 O ⇒ SO_3

b. diphosphorus pentaoxide
2 P and 5 O ⇒ P_2O_5

c. silicon dioxide
1 Si and 2 O ⇒ SiO_2

Common Names for Some Molecular Compounds **(Memorize!)**

H_2O	Water	H_2O_2	Hydrogen Peroxide	O_3	Ozone	CH_4	Methane
C_3H_8	Propane	NH_3	Ammonia	CH_3OH	Methanol	C_2H_5OH	Ethanol
$C_6H_{12}O_6$	Glucose	$C_{12}H_{22}O_{11}$	Sucrose				

Note: Do NOT use prefixes for the above common molecular compounds!

Acid: - ionic substance when dissolves in water will produce an H^+ ion.
- always in aqueous state (*aq*).

$$\boxed{H^+ + \text{Anion} \rightarrow \text{Acid}}$$

Example: $HCl_{(g)} \xrightarrow{H_2O} HCl_{(aq)}$

Nomenclature of Acid

	Ionic Compound Name		Acid Name
1.	hydrogen ~ide	→	hydro~ic acid
2.	hydrogen ~ate	→	~ic acid
3.	hydrogen ~ite	→	~ous acid

Example 3: Name the following acids.

a. $HBr_{(aq)}$

hydrogen bromide

hydrobromic acid

b. $H_2SO_{4\,(aq)}$

hydrogen sulfate

sulfuric acid

c. $H_2SO_{3\,(aq)}$

hydrogen sulfite

sulfurous acid

Example 4: Provide chemical formulas for the following acids.

a. hydrosulfuric acid

hydrosulfuric acid

hydrogen sulfide ⇒ H^+ **and** S^{2-}

$H_2S_{(aq)}$

b. acetic acid

acetic acid ⇒ **hydrogen acetate**

H^+ **and** CH_3COO^-

$HCH_3COO_{(aq)}$

c. hypochlorous acid

hypochlorous acid

hydrogen hypochlorite

⇒ H^+ **and** ClO^-

$HClO_{(aq)}$

6.6: Summary of Naming and Formula Writing

1. Identify whether the compound is ionic, molecular or acid. (Hint: if there is a metal or a complex ion, it is ionic).
2. Molecular compound uses prefixes. They have no charges.
3. Ionic compounds require balancing of charges. Some transition metal cations need to be specified with roman numerals.
4. Acids are originally named with “hydrogen (anion name)”, and they are all in aqueous state. All acids use special rules to name depending on the suffix of the anion.

Assignment

6.5 pg. 159 #37, 38; pg. 160 #39 to 42
6.6 pg. 163 #43 and 44
Ch 6 Review: pg. 166-167 #45 to 71, 73

Chapter 7: Chemical Quantities

7.1: The Mole: A Measurement of Matter

Mole (mol): - a group of atoms or molecules numbered **6.02×10^{23} (Avogadro's Number)**

Example: 1 mol of carbon (C) = 6.02×10^{23} carbon atoms = 12.01 g (same as the amu)
1 mol of oxygen (O_2) = 6.02×10^{23} oxygen molecules = 32.00 g (include subscripts with amu)

Example 1: Calculate the mass of 250 atoms of gold.

1 mol Au = 196.97 g Au = 6.02×10^{23} Au atoms

$$250 \cancel{\text{atoms}} \times \frac{196.97\text{ g}}{6.02 \times 10^{23} \cancel{\text{atoms}}}$$

8.18×10^{-20} g of Au

Example 2: Determine the number of molecules for 50.0 mg of oxygen.

1 mol O_2 = 32.00 g O_2 = 6.02×10^{23} O_2 molecules
(16.00 × 2)

50.0 mg = 0.0500 g

$$0.0500 \cancel{\text{g}} \times \frac{6.02 \times 10^{23}\text{ molecules}}{32.00 \cancel{\text{g}}}$$

9.41×10^{20} molecules of O_2

Molar Mass (g/mol): - the mass per one mole of atoms or molecules.
- sometimes called **gram atomic mass (gam)**, **gram molecular mass (gmm)**, or **gram formula mass (gfm)**.
- molar mass of a mono-atomic element is the same as the atomic mass (gam).
- molar mass of a compound, binary element, or polyatomic element is the same as the combine atomic masses of all atoms in the molecule (gmm).
- molar mass of an ionic compound is the same as the combine atomic masses of all atoms in the ionic formula unit (gfm).

Example 3: Find the molar mass of the following.

a. aluminium

Al = 26.98 g/mol

b. nitrogen

$N_2 = 2 \times 14.01$ g/mol

$N_2 = 28.02$ g/mol

c. water

$H_2O = (2 \times 1.01) + 16.00$

$H_2O = 18.02$ g/mol

d. lead (IV) nitrate

$Pb(NO_3)_4$ = 207.21 + (4 × 14.01) + (12 × 16.00)

$Pb(NO_3)_4$ = 455.25 g/mol

e. sucrose

$C_{12}H_{22}O_{11}$ = (12 × 12.01) + (22 × 1.01) + (11 × 16.00)

$C_{12}H_{22}O_{11}$ = 342.34 g/mol

7.2: Mole-Mass Relationships

Converting between Mass and Moles:

$$\text{Moles (mol)} = \frac{\text{Mass (g)}}{\text{Molar Mass (g/mol)}} \qquad n = \frac{m}{M}$$

n **= moles** ***m*** **= mass** ***M*** **= Molar mass**

Example 1: Calculate the number of moles for:

a. 20.0 g of magnesium chloride

$MgCl_2$ = 24.31 + 2(35.45) M = 95.21 g/mol

$$n = \frac{m}{M} = \frac{20.0\ \cancel{\text{g}}}{95.21\ \cancel{\text{g}}\text{/mol}}$$

n = **0.210 mol**

b. 4.52 mg of glucose

$C_6H_{12}O_6$ = 6(12.01) + 12(1.01) + 6(16.00)

M = 180.18 g/mol

$$n = \frac{m}{M} = \frac{4.52\text{ mg}}{180.18\ \cancel{\text{g}}\text{/mol}}$$

n = **0.0251 mmol**

Example 2: Determine the mass of the following amount.

a. 8.52 mol of ozone

O_3 = 3(16.00) M = 48.00 g/mol

$$n = \frac{m}{M}$$

$m = nM$ = (8.52 $\cancel{\text{mol}}$)(48.00 g/$\cancel{\text{mol}}$)

m = **409 g**

b. 24.7 mmol of phosphoric acid

H_3PO_4 = 3(1.01) + 30.97 +4(16.00) M = 98.00 g/mol

$$n = \frac{m}{M}$$

$m = nM$ = (24.7 m$\cancel{\text{mol}}$)(98.00 g/$\cancel{\text{mol}}$) = 2420.6 mg

m = **2.42 × 10^3 mg = 2.42 g**

Assignment

7.1 pg. 174 #3, 4; pg. 175 #5, 6; pg. 179 #7, 8; pg. 181 #9 to 11, 13, 14
7.2 pg. 183 #16 to 19; pg. 186 #24 and 27

7.3: Percent Composition and Chemical Formulas

Mass Percent: - the mass percentage of each element in a compound.

For Compound $A_xB_yC_z$ with its **Total Mass (*m*)**, the **Mass Percentages** are:

$$\%A = \frac{m_A}{m} \times 100\% \qquad \%B = \frac{m_B}{m} \times 100\% \qquad \%C = \frac{m_C}{m} \times 100\%$$

For Compound $A_xB_yC_z$ with its **Molar Mass (*M*)**, the **Mass Percentages** are:

$$\%A = \frac{(x)(M_A)}{M} \times 100\% \qquad \%B = \frac{(y)(M_B)}{M} \times 100\% \qquad \%C = \frac{(z)(M_C)}{M} \times 100\%$$

Example 1: Propane consists of hydrogen and carbon. If 11.36 g of hydrogen is used to form 62.00 g propane, determine the mass percentages of hydrogen and carbon.

$m_{\text{Total}} = 62.00$ g $\qquad m_H = 11.36$ g

$m_C = 62.00\text{ g} - 11.36\text{ g} \qquad m_C = 50.64$ g

$$\%H = \frac{m_H}{m_{\text{Total}}} \times 100\% \qquad \%C = \frac{m_C}{m_{\text{Total}}} \times 100\%$$

$$\%H = \frac{11.36\text{ g}}{62.00\text{ g}} \times 100\% \qquad \%C = \frac{50.64\text{ g}}{62.00\text{ g}} \times 100\%$$

%H = 18.32 % $\qquad$ **%C = 81.68%**

Example 2: Chromium (II) nitrate has a mass percentage of 29.5% chromium and 15.9% nitrogen. What are the masses for each element in 28.0 g of chromium (II) nitrate.

Chromium (II) nitrate: $Cr(NO_3)_2$ $\qquad m_{\text{Total}} = 28.0$ g $\qquad$ Cr = 29.5% $\qquad$ N = 15.9%

O = 100% – 29.5% – 15.9% $\qquad$ O = 54.6%

$$\frac{m_{Cr}}{m_{\text{Total}}} \times 100\% = \%Cr \qquad \frac{m_N}{m_{\text{Total}}} \times 100\% = \%N \qquad \frac{m_O}{m_{\text{Total}}} \times 100\% = \%O$$

$$m_{Cr} = \frac{\%Cr}{100\%} \times m_{\text{Total}} \qquad m_N = \frac{\%N}{100\%} \times m_{\text{Total}} \qquad m_O = \frac{\%O}{100\%} \times m_{\text{Total}}$$

$$m_{Cr} = 0.295 \times 28.0\text{ g} \qquad m_N = 0.159 \times 28.0\text{ g} \qquad m_O = 0.546 \times 28.0\text{ g}$$

m_{Cr} **= 8.26 g** $\qquad m_N$ **= 1.31 g** $\qquad m_O$ **= 15.3 g**

Example 3: Calculate the mass percentage of sodium chromate.

Na_2CrO_4 $M = 161.98$ g/mol

Assume we have 161.98 g (1 mole) of Na_2CrO_4, there are 2 moles of Na, 1 mole of Cr and 4 moles of O:

$$\% \text{ Na} = \frac{(2\ \cancel{\text{mol}})(22.99\ \cancel{\text{g}}/\cancel{\text{mol}})}{161.98\ \cancel{\text{g}}} \times 100\% = 28.38622052\ \% \qquad \textbf{\% Na = 28.386 \%}$$

$$\% \text{ Cr} = \frac{(1\ \cancel{\text{mol}})(52.00\ \cancel{\text{g}}/\cancel{\text{mol}})}{161.98\ \cancel{\text{g}}} \times 100\% = 32.10272873\ \% \qquad \textbf{\% Cr = 32.103 \%}$$

$$\% \text{ O} = \frac{(4\ \cancel{\text{mol}})(16.00\ \cancel{\text{g}}/\cancel{\text{mol}})}{161.98\ \cancel{\text{g}}} \times 100\% = 39.51105075\ \% \qquad \textbf{\% O = 39.511 \%}$$

Empirical Formula: - the simplest ratio between the elements in a chemical formula.

Molecular Formula: - the actual chemical formula of a compound.

Molecular Formula = (Empirical Formula)$_n$

where n = natural number

Example 4: $C_6H_{12}O_6$ $\longleftarrow$ CH_2O

Molecular Formula for Glucose — Empirical Formula

Note: Knowing the mass % of a compound allows us to find the empirical formula. To know the molecular formula, we must also know the molar mass.

Example 5: Vitamin C has a molar mass of 176.14 g/mol and contains carbon, hydrogen, and oxygen atoms. If the % mass of carbon and oxygen are 40.91% and 54.50% respectively, determine the empirical and molecular formula of vitamin C.

% C = 40.91%   % O = 54.50%   % H = 100% − 40.91% − 54.50% = 4.59%

Assume 100 g of Vitamin C. Then, there are

m_C = 100 g × 40.91% = 40.91 g m_O = 100 g × 54.50% = 54.50 g m_H = 100 g × 4.59% = 4.59 g

$$n_C = \frac{40.91\,\cancel{g}}{12.011\,\cancel{g}/\text{mol}} = 3.406044459\ \text{mol}_C \qquad n_H = \frac{4.59\,\cancel{g}}{1.0079\,\cancel{g}/\text{mol}} = 4.554023217\ \text{mol}_H$$

$$n_O = \frac{54.50\,\cancel{g}}{16.00\,\cancel{g}/\text{mol}} = 3.40625\ \text{mol}_O$$

$$\frac{n_C}{n_O} = \frac{3.406044459\ \text{mol C}}{3.40625\ \text{mol O}} \approx \frac{1\ \text{mol C}}{1\ \text{mol O}} \qquad \frac{n_H}{n_O} = \frac{4.554023217\ \text{mol H}}{3.40625\ \text{mol O}} \approx 1.33 = \frac{4\ \text{mol H}}{3\ \text{mol O}}$$

$n_C : n_O = 1 : 1$ ⟵ **Combine Ratios** ⟶ $n_H : n_O = 4 : 3$

Empirical Formula = $C_3H_4O_3$ **(88.07 g/mol)**

$$\frac{\text{Actual Molar Mass}}{\text{Emprical Molar Mass}} = \frac{176.14\,\cancel{g/mol}}{88.07\,\cancel{g/mol}} = 2$$

Molecular Formula = Empirical Formula × 2

Molecular Formula = $C_6H_8O_6$

OR

Another Method may be used where the Actual Molar Mass becomes the Mass of Vitamin used. Then, the Mole of each Atom is calculated to determine the Molecular Formula first.

$$n_C = \frac{40.91\% \times 176.14\,\cancel{g}}{12.011\,\cancel{g}/\text{mol}} \approx 6.00\ \text{mol}_C \qquad n_H = \frac{4.59\% \times 176.14\,\cancel{g}}{1.0079\,\cancel{g}/\text{mol}} \approx 8.00\ \text{mol}_H$$

$$n_O = \frac{54.50\% \times 176.14\,\cancel{g}}{16.00\,\cancel{g}/\text{mol}} \approx 6.00\ \text{mol}_O$$

Molecular Formula ($C_6H_8O_6$) will be found first, then the Empirical Formula ($C_3H_4O_3$) will be stated.

Assignment

7.3 pg. 189 #29, 30; pg. 191 #31 to 33; pg. 192 #34; pg. 193 # 35, 36; pg. 194 #37, 38; pg. 195 #39 to 43

Ch 7 Review: pg. 198 #45 to 52, 55, 56, 60 to 66

Assignment

Unit 1 Review: pg. 81 #1 to 12; pg. 105 #1 to 15; pg. 49 #1 to 11, 15 to 17; pg. 131 #1 to 9; pg. 169 #1 to 15; pg. 201 #1 to 3, 4 to 14