Unit 1: BASIC CHEMISTRY

Chapter 1: Introduction

1.2: The Scientific Method

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

Steps of Scientific Method

1. **Observation**: - an act of recognizing and noting a fact or occurrence.
2. **Scientific Hypothesis**: - an **educated guess** or a ***testable*** **assumption** to explain any observable phenomenon.
3. **Experimentation** or **Control Test**: - a test performed by scientist and researcher to increase the accuracy and reliability of an experimental test.
   - involves testing of two variables **(manipulated and responding variables)** while all other variables are controlled.
   a. **Accuracy**: - sometimes refer to as **validity**. It describes whether the result is correct.
   b. **Reproducibility**: - sometimes refer to as **reliability**. It describes the consistency and the repeatability of the result.
   c. **Observations**: - often involve making measurements with scientific instrument(s).
4. **Theory**: - an idea that can explain a set of observations that has stood up to repeated scrutiny.
5. **Scientific Law**: - a concise statement that summaries the results of many observations and experiments.
   - describes the phenomenon without trying to explain it.

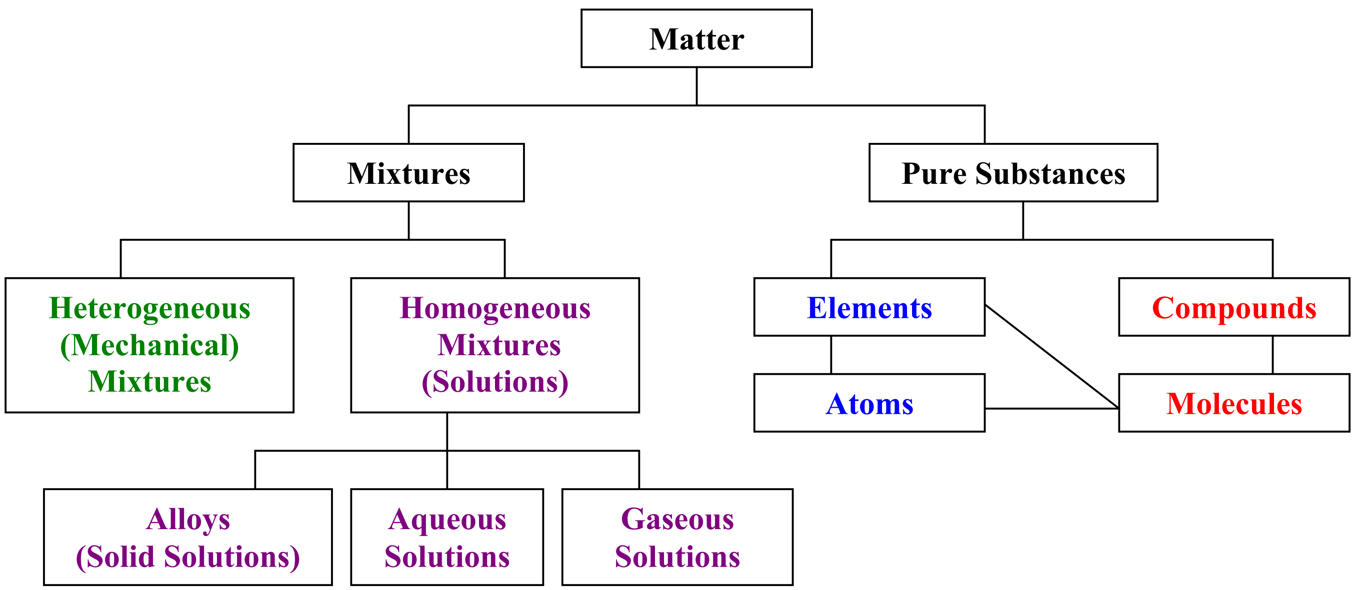
Limitations of Science: - science cannot answer all questions. It can only tackle "testable" hypothesis. ***Philosophical and Religious Questions CANNOT be answered by science.***

1.3: Classification of Matter

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

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

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

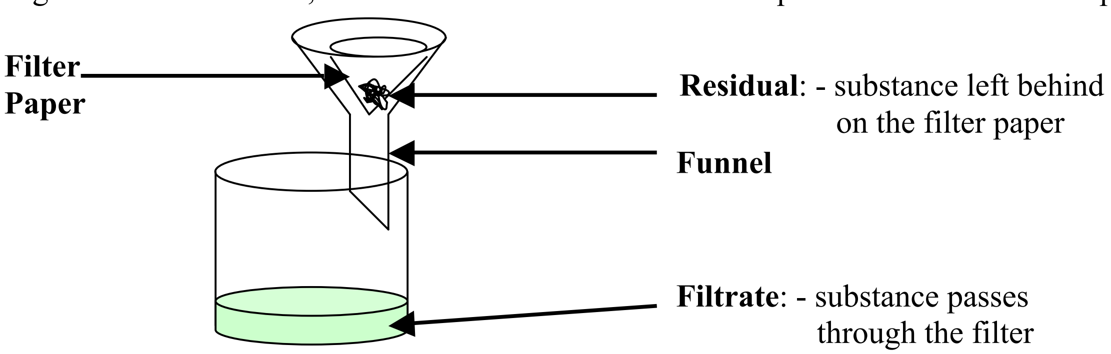


Matter: - anything that has a mass and occupies space.

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 liquids and solids can be separated.



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

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

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

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

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

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

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

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

Atom: - the smallest particle of matter.

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

Examples:

1. Classify the following as Heterogeneous or Homogeneous Mixture:
   a) a bag of gravel b) cement c) saturated salt water d) a methanol and water mixture
   e) oil and water f) the atmosphere g) Jell-O h) diet carbonated soft drink
2. Classify the following as Mixture or Pure Substance:
   a) lake water b) tap water c) distilled water d) iron e) steel (iron and carbon)
   f) chromium g) beer h) sugar i) gasoline
3. Classify the following as Element or Compound: (use the Periodic Table of Elements)

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

Answers:

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

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

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

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

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

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

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

1.4: Physical and Chemical Properties of Matter

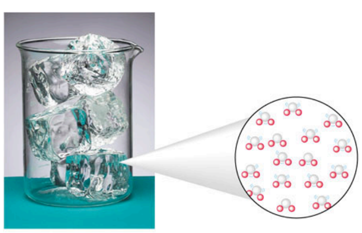
Physical Property: - an observation or measurement that does **not** change the composition or identity of substance.

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

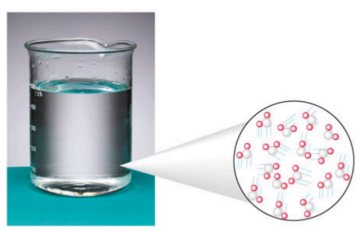
Liquid: - the state of matter where it has a definite volume but an indefinite shape.

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

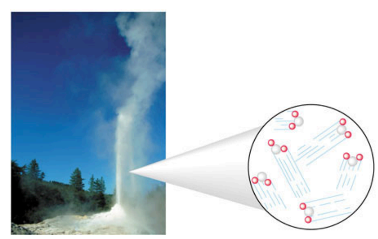
Kinetic Molecular Theory and States of Matter



Solids have particles are in fixed positions.

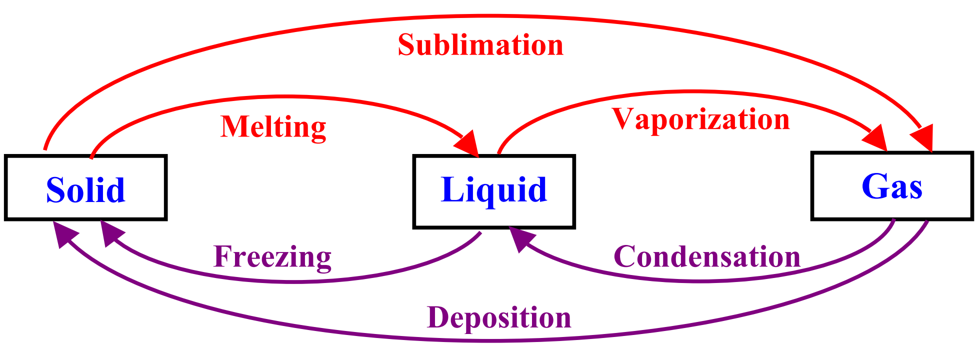


Liquids have particles that can "rolled" past one another.



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

Phase Change



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

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

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

Measurable Property: - a quantitative observation or measurement.

a) **Extensive Property**: - a measurement that depends on how much matter is being considered. (Examples: mass and volume)

b) **Intensive Property**: - does **not** depend on how much matter is being considered. (Example: density – both mass and volume change proportionally)

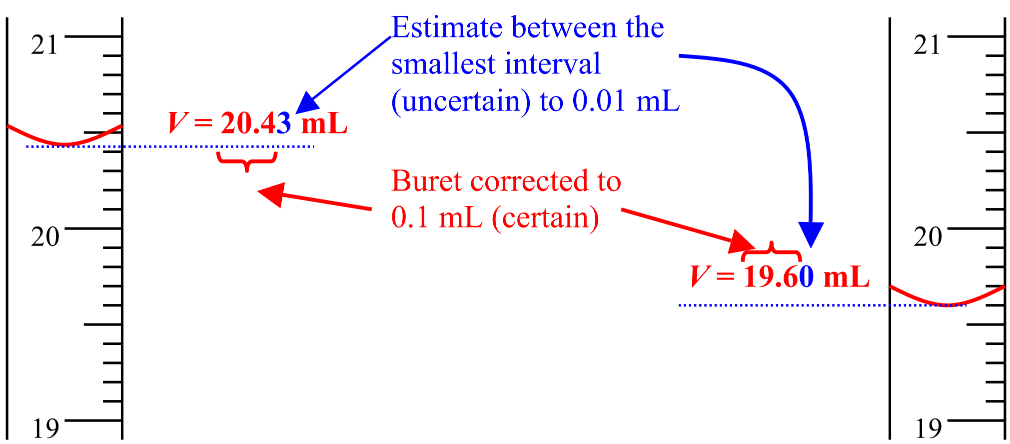
Assignment
1.3 & 1.4 pg. 23 #2 to 12 (even)

1.5 & 1.6: Measurement & Handling Numbers

Macroscopic Properties: - can be measured directly with instruments like scales and graduated cylinders, and thermometers.

Microscopic Properties: - must be determined by indirect methods like mass spectrometers; linear accelerators.

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



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

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

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

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

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

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

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

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

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

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

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

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

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

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

To Count Significant Digits

1. Start counting the first non-zero digit. Do NOT count the leading zero(s).
2. Count all **captive zeros** (between non-zero digits) and **trailing zero** at the end of the measurement.
3. Include ALL digits of a whole number ***if*** it contains a decimal point. All digits in a whole number that does not end in zero(s) are significant even if it does not contain a decimal point. (Examples: 420. → 3 sig digs; 402 or 402. → 3 digits)
4. If a measurement **contains no decimal places**, the **trailing zeros may or may not be significant**. Hence, using scientific notation eliminates this ambiguity.

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

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

Calculating with Significant Digit

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

Example 3: 5.345 g + 0.42 g + 11.8 g

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

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

17.6 g

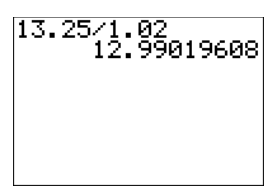
2. **Multiplying and Dividing**: - answer should be in the least number of significant digits used in calculation.

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

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

The least number of significant digits used is three.

13.0 g/mL



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

   (Note: most chemistry textbooks round off at every step. In essence, the answers vary little. Hence, either way is considered correct.)

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

$$E_{output} = mc^2 + E_{input}$$
$$E_{output} = (5.3 \times 10^{-3}\text{ kg})(3.00 \times 10^8\text{ m/s})^2 + 4.15 \times 10^{14}\text{ J}$$
$$E_{output} = 8.92 \times 10^{14}\text{ J}$$
$$\mathbf{E_{output} = 8.9 \times 10^{14}\text{ J}}$$

Scientific Notation ($\times 10^n$)

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

2nd

EE

,

Theoretical Result: - the supposed result of an experiment 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%**

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 cm³

$\times$ 1000 $\times$ 1000

1 L = 1000 cm³

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 7: 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 × 10^4** mg **(right 6 places)**

two significant digits

Temperature: - the average kinetic energy of a substance.

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

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

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

Convert from degree Celsius to Fahrenheit

32 is an exact number

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

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

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

$T_F = -34.6$ F

$T_F = -34.6$ F

$T_K = T_C + 273.15$

$T_K = -37.0 + 273.15$

$T_K = 236.15$ K

$T_K = 236$ K

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.

Density: - the amount of mass per unit of volume

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

Example 9: 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:

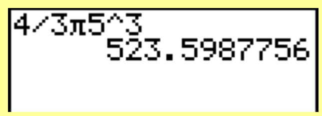
$$D = \frac{m}{V}$$
$$DV = m$$

We need to use the Volume formula of a Sphere.

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

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

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

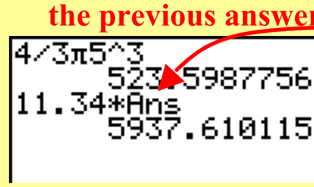


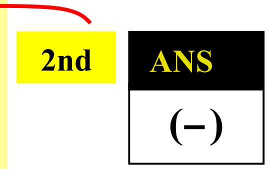
(Do NOT round off. We are not done yet.)

Substitute D and V to solve for m

To recall all digits of the previous answer

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





Assignment

1.5 & 1.6 pg. 23−24 #14, 16, 18 to 20, 22, 24 to 26, 28 to 30

1.7: Dimensional Analysis in Solving Problems

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

Example 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 or 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 or 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 or 23 km/L

Assignment

1.7 pg. 24 #32 to 34, 36, 38, 40

Chapter 2: Atoms, Molecules and Ions

2.1: The Atomic Theory

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 and have the same physical and chemical properties. Different elements have different kind of atoms.**
3. **Chemical compounds are formed when different kinds of atoms combine together. A particular compound always has the same relative numbers and types of atoms.**
4. **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 (they cannot be created or destroyed) in a chemical reaction.**

Early Fundamental Chemical Laws

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

Example 1: A 7.00 g of iron nail is allowed to rust. The rusted nail has a mass of 10.2 g. What is the amount of oxygen reacted with the iron nail?

Iron + Oxygen → Rust
7.00 g ? g 10.2 g

Mass of Oxygen = 10.2 g – 7.00 g

Mass of Oxygen = 3.2 g

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

Example 2: Water contains about 8 parts oxygen to 1 part hydrogen by mass. A 192 g of unknown liquid compose 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.

3. **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 3: State the ratios of carbon between the following hydrocarbon compounds.

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

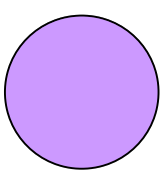
$$\frac{A}{B} = \frac{2.973\text{ g}}{3.963\text{ g}} = 0.7501892506 \approx 0.75 \qquad \frac{\boldsymbol{A}}{\boldsymbol{B}} = \frac{\mathbf{3}}{\mathbf{4}}$$

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

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

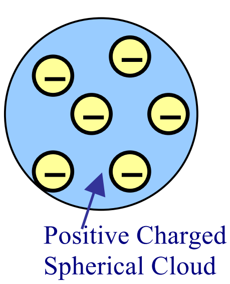
2.2: The Structure of the Atom

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.



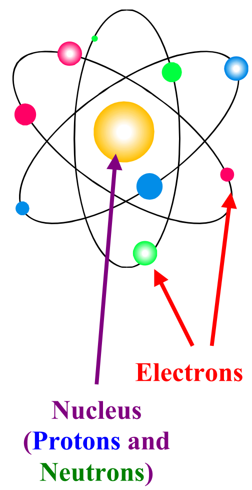
Dalton's Atomic Model:

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. He also found the charge to mass ratio of an electron to be -1.76×10^8 C/kg. In 1917, 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.

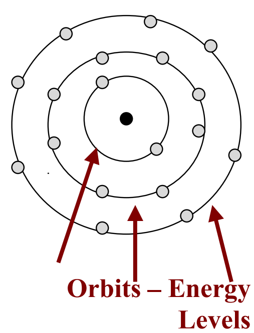


Nuclear Model:

In 1895, Wilhelm **Röntgen** found that cathode rays caused glass and metals to emit unusual rays (he called them X-rays). Later, Antoine **Becquerel** discovered that a piece of radioactive metal (uranium) can expose a photographic plate. A student of Becquerel, **Marie Curie** termed this energy as radiation. Three types of rays were found as they deflected differently through a magnetic field. **Alpha (α) rays** – positively charged particles, **Beta (β) rays** – negatively charged particles, and **Gamma (γ) rays** – neutrally charged particles with no deflection in a magnetic field.

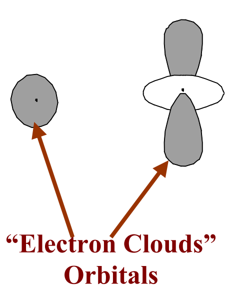
In 1912, Ernest **Rutherford** proposed the Nuclear Model for atoms after his famous **gold foil experiment** (he shot alpha particles into a piece of gold foil and found that the alpha particles passed through the gold foil – indicating the the atom is made up of mostly empty space). Earlier to this 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.

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

2.3: Atomic Number, Mass Number and Isotopes

Assignment
2.2 pg. 53 #4, 8

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 **mass number**. This is the mass characteristic of a given element**. The mass number 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. Mass Number has a unit of amu (Atomic Mass Unit).

Because different elements have different mass number, 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 mass number**, 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 mass number (rounded off whole number) with the atomic number**. (*Note*: The Table of Elements usually shows average atomic mass – more on this in the next chapter).

Atomic Number = Number of Protons and Electrons of an Atom

Number of Neutrons = Mass Number – Atomic Number

Example 1: State the Atomic Number, Average 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^-**) **Average Atomic Mass = 35.45** # of Neutron = 35.45 – 17 = **18 n**	1 1.01 **H** **Hydrogen**	**Atomic Number = 1** (**1 p^+** and **1 e^-**) **Average 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 → A
$${}^{A}_{Z}X$$ ← Element Symbol
Atomic Number → Z

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

Using the same rule, the energy level diagram of an isotope can be drawn as well. Recall that the superscript is the atomic mass of the isotope and the subscript is the atomic number. Since the atomic mass of the isotope is different than the atomic mass of the original element, the number of neutrons of the isotope is different than the number of neutrons of the element.

Electron Shells, Energy Levels and Valence Electrons:

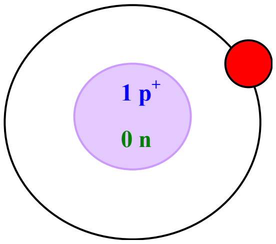
Recall from the Bohr Model studied earlier. It states that **electrons travel around the nucleus in specified orbits (electrons are quantum)**. These orbits are called **energy levels**. They can also be called **electron shells**. These orbits are very similar to the planets orbiting our sun. The only 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 2: 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.

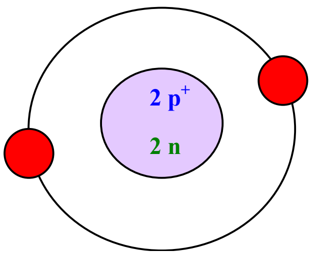


Element: Hydrogen (H)

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

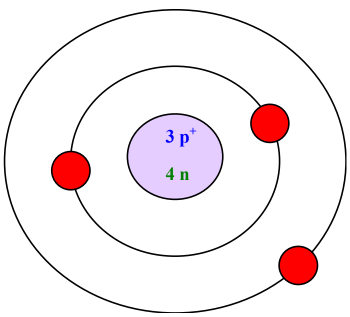


Element: Helium (He)

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

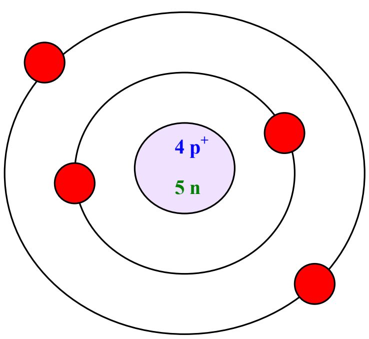


Element: Lithium (Li)

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

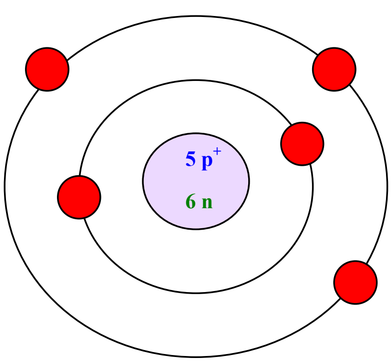


Element: Beryllium (Be)

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

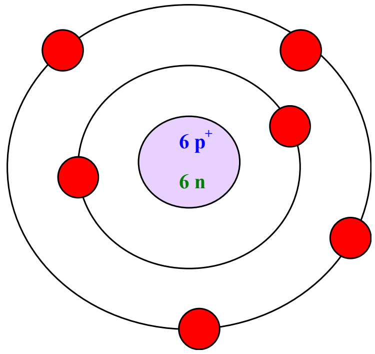


Element: Boron (B)

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

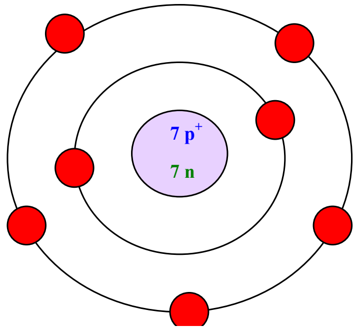


Element: Carbon (C)

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

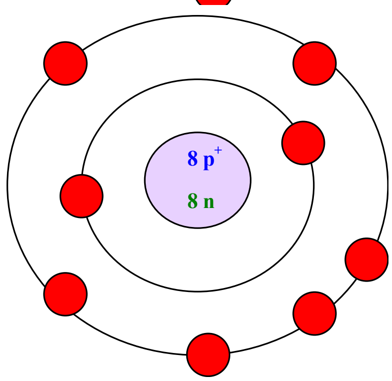


Element: Nitrogen (N)

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

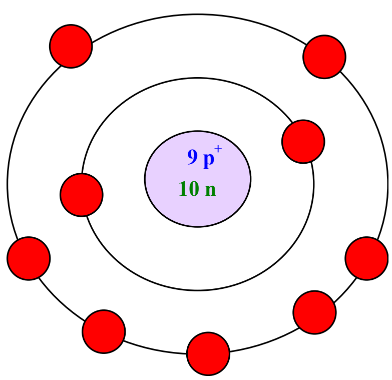


Element: Oxygen (O)

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

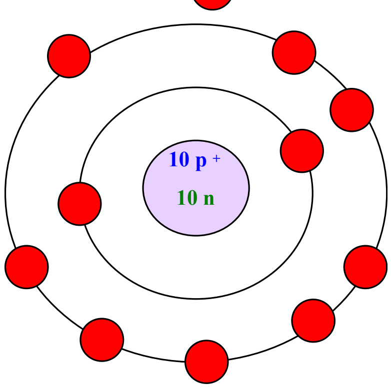


Element: Fluorine (F)

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)

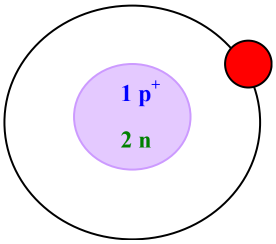
Atomic Number: 10 **Atomic Mass: 20.18** **Nucleus: 10 p^+ and 10 n**

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.

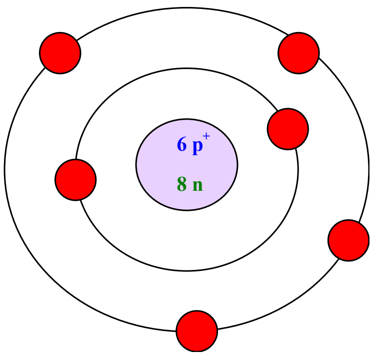
Example 3: Draw the Bohr Energy Level diagram for tritium, ${}^{3}_{1}H$ and carbon-14. State the number of protons, electrons, and neutrons as well as the location for each of these elements.



Isotope: Tritium (${}^{3}_{1}H$)

Atomic Number: 1 **Atomic Mass: 3** **Nucleus: 1 p^+ and 2 n**

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



Isoptope: Carbon-14 (${}^{14}_{6}C$)

Atomic Number: 6 **Atomic Mass: 14.01** **Nucleus: 6 p^+ and 8 n**

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

2.5: Molecules and Ions

Assignment
2.3 pg. 54 #10 to 16 (even)

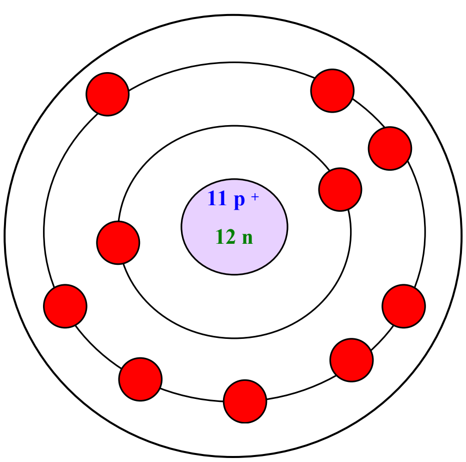
Molecule: - basic unit of a compound.
- contains at least two atoms of the same or different kind of elements.

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

1. **Cations**: - **positive** charged ions (atoms that lose electrons).
   - to name cations → element name follows by the word "ion"

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

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



Sodium Ion (Na^+)

Atomic Number: 11 **Atomic Mass: 22.99**

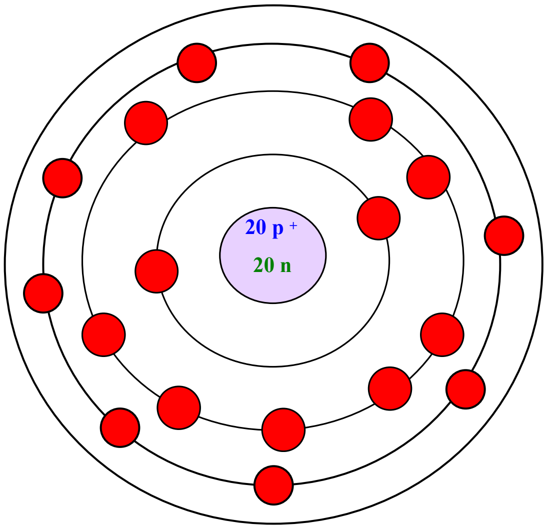
Nucleus: 11 p^+ and 12 n

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

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

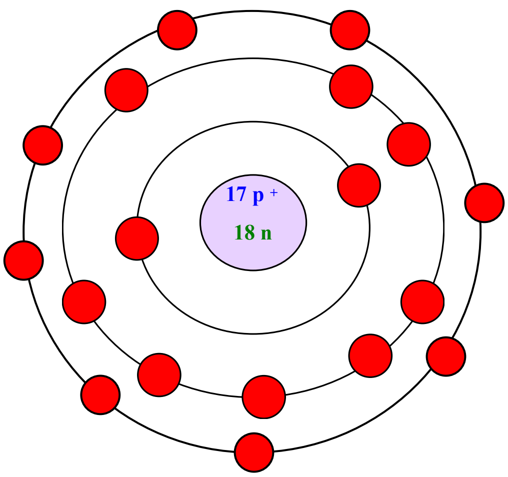
Location on the Period Table of Elements:

Fourth Row; Column IIA

2. **Anions**: - **negative** charged ions (atoms that gain electrons).
- to name anions → keep the first part of element name follow by suffix ~*ide*

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

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



Chloride (Cl^-)

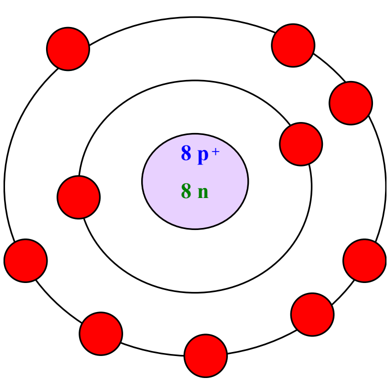
Atomic Number: 17 **Atomic Mass: 35.45**
Nucleus: 17 p^+ and 18 n

3rd Energy Level: 8 e^- (8 valence e^- - Filled)
2nd Energy Level: 8 e^-
1st 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

2nd Energy Level: 8 e^- (8 valence e^- - Filled)
1st 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 electrons – for main groups – IA to VIIIA columns) to achieve stability.
- in most cases, this means having the same electron arrangement of the **nearest** noble gas.
- exceptions to the rule include helium (only 2 electrons to fill the first energy level), and the transition metals.

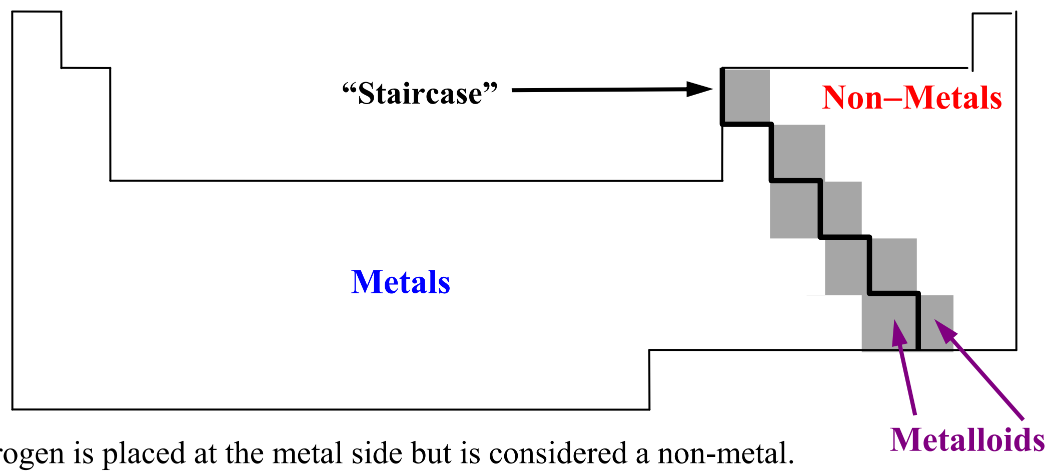
Assignment
2.5 pg. 54−55 #26 to 34 (even)

2.4: The Periodic Table

Mendeleev's Periodic Table of Elements: - Dmitri Mendeleev organized the elements on a table based on the atomic masses as well as the "recurring" chemical and physical properties.
- the columns of elements with the same chemical and physical properties are called groups or families.
- the rows of the elements where they exhibit a gradual change in chemical properties are called periods.

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 conductors of heat and electricity.
5. Metals are lustrous (shiny).

Physical Properties of Non-Metals:

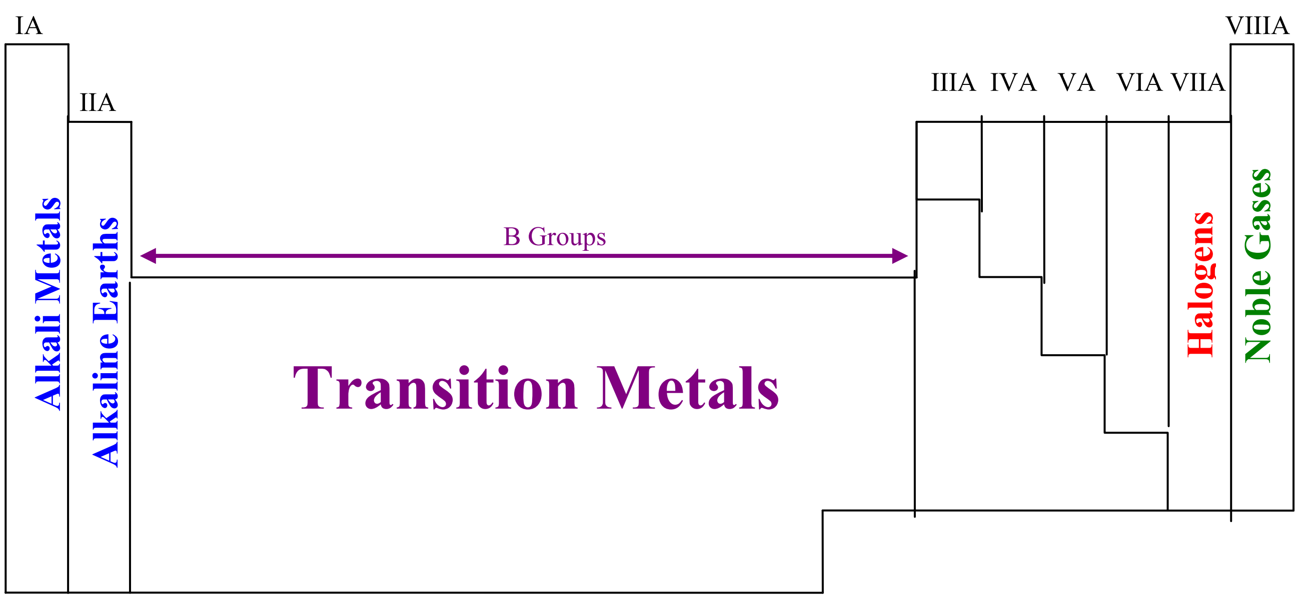
1. Non-metals are mostly gases or solids at room temperature (with the exception of bromide).
2. Non-metals usually do not have the physical properties of metals.

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.

Today, all elements and their symbols are listed in an orderly fashion with the help of the Periodic Table of Elements. The use of standardized symbols allows scientists from all over the world to share their knowledge despite the differences in language. Most elements are **monoatomic**. That means their atoms can exist individually ("*mono*" means one). Others are **diatomic**, atoms that exist in pairs ("*di*" means two). Some are **polyatomic**, atoms that exist in numbers more than one ("*poly*" means many). The table below shows all the diatomic and polyatomic elements.

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)	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.** Most symbols are recognizable from the name of the elements (Zinc: Zn; Carbon: C; Aluminium: Al). Others 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 familiarize themselves with the whereabouts of the elements on the Table.

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

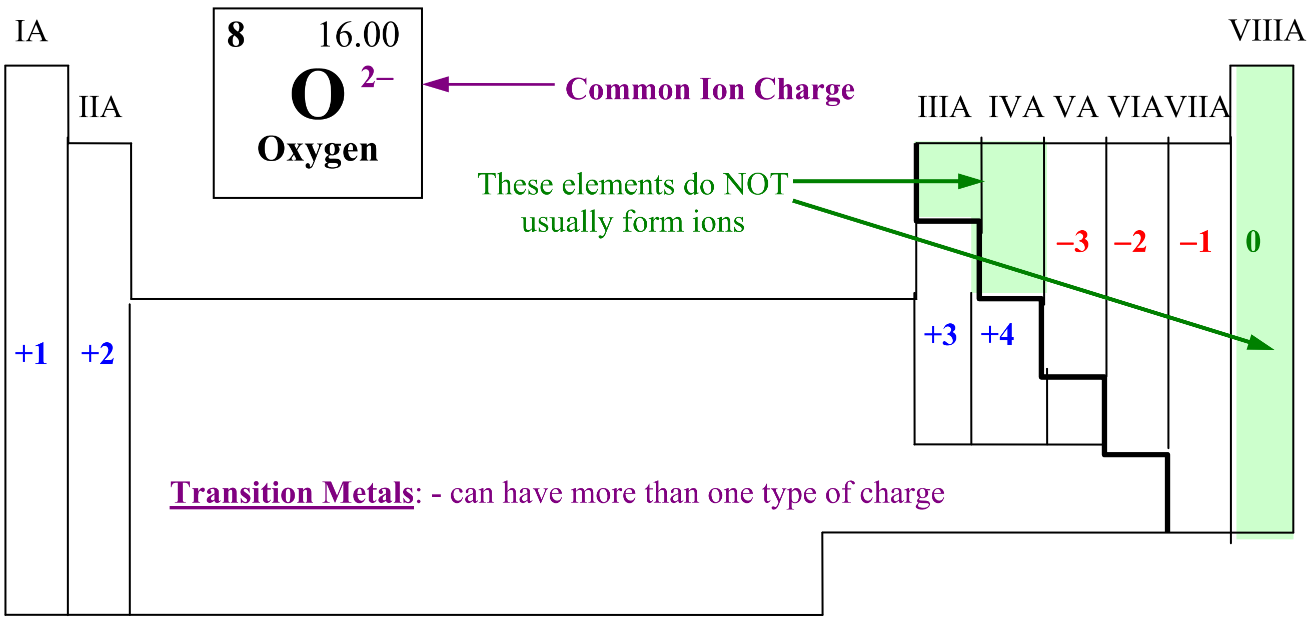
26	55.85
	3+, 2+
Fe	
Iron	

Fe^{2+} = Iron (II) ion (Old Name – Ferr*ous* ion)

Fe^{3+} = Iron (III) ion (Old Name – Ferr*ic* ion)
(First Charge Listed is the most common type)

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



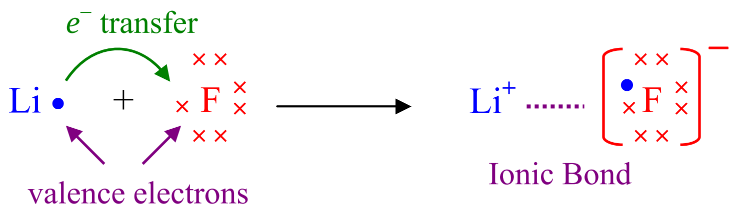
The reason that noble gases (column VIIIA) do not form ions is because their outermost shells are filled with the maximum number of electrons allowed. That is why we call this group of elements "noble gases". They do not form ions because they are stable. Hence we use the word "noble" to describe them. 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 summarizes 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

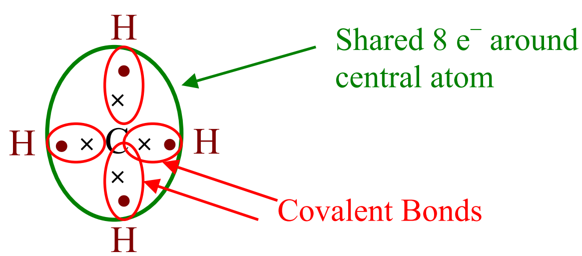
Ionic Compound: - when a metal element combines with a non-metal element.
- forms **ionic bonds** (electrons are "stolen" or "transferred" from one atom to another).
- dissociates into **electrolytes** (forms ions, cations and anions, when dissolve in water).

Example: LiF



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

Example: CH_4



Monoatomic Ions: - ions that came from a single atom (include metal cations and non-metal anions).
- monoatomic anion ends with suffix ~ide.
- some transition metal ions require Roman Numeral when there can be more than one type of charge.

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

Polyatomic 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

Assignment
2.4 pg. 54 #18 to 24 (even)

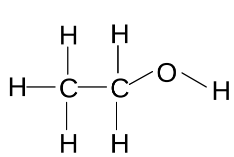
2.6: Chemical Formulas

Chemical Formulas: - using atomic symbols, they express the compositions of molecular and ionic compounds.

Allotropes: - different distinct forms of an element. (Example: Carbon as diamond or graphite)

Structural Formula: - a diagram that shows how the atoms in a molecule are bonded and arranged

Example: Ethanol has a chemical formula C_2H_5OH.



(Note that we do not combine the H in the chemical formula because we want to emphasize the molecular structure.)

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: $C_6H_{12}O_6$ ⟵ CH_2O

Molecular Formula for Glucose — Empirical Formula

2.7A: Naming Simple Ionic Compounds

Organic Compounds: - sometimes called hydrocarbons which contain carbon and hydrogen, and may also contain oxygen, nitrogen and sulphur.

Inorganic Compounds: - compounds that are not hydrocarbons.

Nomenclature: - a naming system.

IUPAC: - International Union of Pure and Applied Chemistry.
- an organization that oversees the standard regarding chemistry including chemical nomenclature.

Nomenclature of Ionic Compounds

1. **Balance the Cation and Anion Charges.**
2. Use **brackets** for **multiple 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 formula of the followings.

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

Oxoanions: - 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 oxoanions.

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.
- to name hydrates → use the ionic compound name, then write the 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

2.7B: Naming Simple 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: - an ionic substance when dissolves in water produces an H^+ ion.
- chemical formula usually starts with H.
- 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 brom*ide*

hydro*brom*ic acid

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

hydrogen sulf*ate*

sulfur*ic acid*

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

hydrogen sulf*ite*

sulfur*ous acid*

Example 4: Provide chemical formula for the following acids.

a. hydrosulfuric acid b. acetic acid c. hypochlorous acid

hydro**sulfur*****ic acid***

hydrogen sulf***ide*** ⇒ **H^+ and S^{2-}**

$H_2S_{(aq)}$

acet***ic acid*** ⇒ **hydrogen acet*****ate***

H^+ and CH_3COO^-

$HCH_3COO_{(aq)}$

hypochlor***ous acid***

hydrogen hypochlor***ite***

⇒ **H^+ and ClO^-**

$HClO_{(aq)}$

<u>Oxoacids</u>: - acids that consist of oxoanions.

Examples: $HBrO_{4\,(aq)}$ – perbromic acid; $HIO_{2\,(aq)}$ – iodous acid

<u>Summary of Naming and Formula Writing</u>

1. Identify whether the compound is ionic, molecular or acid. (Hint: if there is a metal or a polyatomic 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 "hydrogen - anion name. They are all in aqueous state. All acids use special naming rules depending on the suffix of the anion.

<u>Assignment</u>

2.6 pg. 55 #36, 38, 42, 44, 46

2.7 pg. 55 #47 to 50

Chapter 3: Stoichiometry

3.1: Atomic Mass

Atomic Mass: - sometimes called **atomic weight**.
- the mass of the atom in atomic mass unit (amu).
- 1 amu = exactly one-twelfth the mass of one carbon-12 atom ≈ 1.67×10^{-27} kg.

Average Atomic Mass: - Average Mass of an atom and its isotopes after accounting their proportions of abundance (as stated on the Periodic Table of Elements).

Relative Abundance: - the relative proportion of various isotopes of an element.

Example 1: 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 ${}^{3}_{1}\text{H}$.

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

Average amu of Hydrogen = 1.008

Example 2: Lithium has two naturally occurring isotopes. Lithium-6 has an atomic mass of 6.015 amu; lithium-7 has an atomic mass of 7.016 amu. The average atomic mass of lithium is 6.941 amu. Determine whether lithium-6 or lithium-7 is the naturally occurring atom and their percentages of abundance.

Let x = the proportion of lithium-6
Then $(1 - x)$ = proportion of lithium-7

Note: Since the average atomic mass of lithium is 6.941, which is closer to the amu of lithium-7, we can assume that the most abundant isotope is lithium-7.

Average amu of Lithium = $(x)(6.015) + (1 - x)(7.016)$

$6.941 = 6.015x + 7.016 - 7.016x$ (Multiply into binomial)

$6.941 - 7.016 = 6.015x - 7.016x$ (Combine Like-Terms)

$-0.075 = -1.001x$ (Solve for x)

$$\frac{-0.075}{-1.001} = x$$

$x = 0.07492507$ (Need to round to 4 significant digits)

Lithium-6 occurs 7.493%.

Lithium-7 occurs (100% − 7.493%) = 92.51%

3.2: Avogadro's Number and the Molar Mass of an Element

Mole (mol): - a group of atoms or molecules numbered **6.022×10^{23} (*Avogadro's Number, N_A*)**

Examples: 1 mol of carbon (C) = 6.022×10^{23} carbon atoms = 12.01 g (same as the amu)
1 mol of oxygen (O_2) = 6.022×10^{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 = 6.022×10^{23} Au atoms

$$250 \cancel{\text{atoms}} \times \frac{196.97\text{ g}}{6.022\times10^{23}\cancel{\text{ atoms}}}$$

8.177×10^{-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 = 6.022×10^{23} O_2 molecules
(16.00 × 2)

50.0 mg = 0.0500 g

$$0.0500 \cancel{\text{g}} \times \frac{6.022\times10^{23}\text{ molecules}}{32.00\cancel{\text{ g}}}$$

9.41×10^{20} molecules of O_2

3.3: Molecular Mass

Molar Mass (g/mol): - sometimes refer to as **molecular mass** or **molecular weight**, is the mass per one mole of atoms or molecules.
- molar mass of a mono-atomic element is the same as the atomic mass.
- molar mass of a compound, binary element, or polyatomic element is the same as the combine atomic masses of all atoms in the molecule.

Example 1: 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

$H_2O = (2 \times 1.008) + 16.00$

H_2O = 18.02 g/mol

d. lead (IV) nitrate

$Pb(NO_3)_4 = 207.2 + (4 \times 14.01) + (12 \times 16.00)$

$Pb(NO_3)_4$ = 455.2 g/mol

e. sucrose

$C_{12}H_{22}O_{11} = (12 \times 12.01) + (22 \times 1.008) + (11 \times 16.00)$

$C_{12}H_{22}O_{11}$ = 342.3 g/mol

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 2: Calculate the number of moles for

a. 20.0 g of magnesium chloride

$MgCl_2 = 24.30 + 2(35.45)$ $M = 95.20$ g/mol

$n = \dfrac{m}{M} = \dfrac{20.0\ \cancel{\text{g}}}{95.20\ \cancel{\text{g}}/\text{mol}}$ **$n = 0.210$ mol**

b. 4.52 mg of glucose

$C_6H_{12}O_6 = 6(12.01) + 12(1.008) + 6(16.00)$

$M = 180.2$ g/mol

$n = \dfrac{m}{M} = \dfrac{4.52\ \text{m}\cancel{\text{g}}}{180.2\ \cancel{\text{g}}/\text{mol}}$ **$n = 0.0251$ mmol**

Example 3: Determine the mass of the following amount.

a. 8.52 mol of ozone

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

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

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

$m = 409$ g

b. 24.7 mmol of phosphoric acid

$H_3PO_4 = 3(1.008) + 30.97 + 4(16.00)$ $M = 97.99$ g/mol

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

$m = nM = (24.7\ \text{m}\cancel{\text{mol}})(97.99\ \text{g}/\cancel{\text{mol}}) = 2420.353$ mg

$m = 2.42 \times 10^3$ mg $= 2.42$ g

3.4: The Mass Spectrometer

Mass Spectrometer: - an instrument that measures the relative abundance of an element.
- uses magnetic and electric fields to deflect different charged isotopes. The heavier the isotope, the bigger the radius of the particle (deflected least).
- the result can be graphed into a mass spectrum.

$$F_{\text{centripetal}} = F_{\text{magnetic}}$$

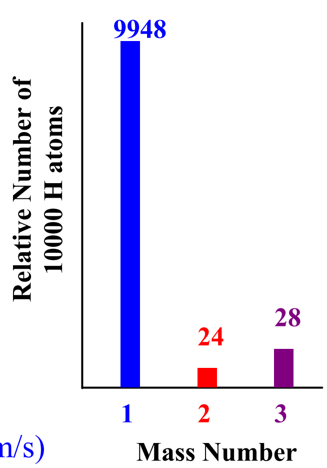
$$\frac{mv^2}{r} = qvB$$

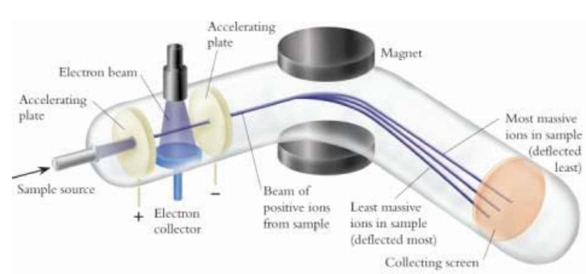
$$m = \frac{qvBr}{v^2}$$

$$m = \frac{qBr}{v}$$

$$m \propto r$$

q = Charge in Columb (C)
v = speed of charged particle (m/s)
B = magnetic field in Tesla (T)
r = radius of deflection (m)





Schematic representation of one type of mass spectrometer. An electron beam fragments gas atoms or molecules into positively charged ions. The ions are accelerated and then deflected by a magnet. More massive particles are deflected by smaller amounts than less massive particles.

Assignment

3.1 pg. 86 #4, 6, 8
3.2 pg. 86–87 #10, 13 to 16, 18 to 20, 22
3.3 pg. 87 #24 to 30 (even)

3.5 & 3.6: Percent Composition of Compounds & Experimental Determination of Empirical Formulas

Mass Percent: - also called percent composition by mass.
- it is 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: Calculate the mass percentage of sodium chromate.

Na_2CrO_4 $\quad M = 161.98$ 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:

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

% Na = 28.39 %

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

% Cr = 32.10 %

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

% O = 39.51 %

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: $C_6H_{12}O_6 \longleftarrow CH_2O$

Molecular Formula for Glucose — Empirical Formula

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

Example 1: 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

$$n_C = \frac{40.91\,\cancel{g}}{12.01\,\cancel{g}/\text{mol}} = 3.40632806\ \text{mol}_C \qquad n_H = \frac{4.59\,\cancel{g}}{1.008\,\cancel{g}/\text{mol}} = 4.553571429\ \text{mol}_H$$

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

$$\frac{n_C}{n_O} = \frac{3.40632806\ \text{mol C}}{3.40625\ \text{mol O}} \approx \frac{1\ \text{mol C}}{1\ \text{mol O}} \qquad \frac{n_H}{n_O} = \frac{4.553571429\ \text{mol H}}{3.40625\ \text{mol O}} \approx 1.33 = \frac{4\ \text{mol H}}{3\ \text{mol O}}$$

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

Empirical Formula = $C_3H_4O_3$ **(88.06 g/mol)**

$$\frac{\text{Actual Molar Mass}}{\text{Emprical Molar Mass}} = \frac{176.14\ \text{g}/\cancel{\text{mol}}}{88.06\ \text{g}/\cancel{\text{mol}}} = 2$$

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.

$$n_C = \frac{40.91\% \times 176.14\,\cancel{g}}{12.01\,\cancel{g}/\text{mol}} \approx 6.00\ \text{mol}_C \qquad n_H = \frac{4.59\% \times 176.14\,\cancel{g}}{1.008\,\cancel{g}/\text{mol}} \approx 8.00\ \text{mol}_H$$

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

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

Example 2: Octane, which contains carbon and hydrogen, is burned to produce 4.40 g of carbon dioxide gas and 2.03 g of water vapour. Determine the empirical and molecular formula of octane if it is found to have a molar mass of 114.26 g/mol from an analysis using a mass spectrometer.

$$n = \frac{4.40\,\cancel{g}}{44.01\,\cancel{g}/\text{mol}} = 0.0999772779\ \text{mol}\ CO_2 \qquad n = \frac{2.03\,\cancel{g}}{18.02\,\cancel{g}/\text{mol}} = 0.1126526082\ \text{mol}\ H_2O$$

$n_C = 0.0999772779\ \text{mol}_C$ $n_H = 0.1126526082\ \text{mol} \times 2 = 0.2253052164\ \text{mol}_H$

$$\frac{n_H}{n_C} = \frac{0.2253052164\ \text{mol H}}{0.0999772779\ \text{mol C}} \approx \frac{2.25\ \text{mol H}}{1\ \text{mol C}} = \frac{9\ \text{mol H}}{4\ \text{mol C}}$$

Empirical Formula = C_4H_9 **(57.11 g/mol)**

$$\frac{\text{Actual Molar Mass}}{\text{Emprical Molar Mass}} = \frac{114.26\ \text{g}/\cancel{\text{mol}}}{57.11\ \text{g}/\cancel{\text{mol}}} \approx 2$$

Molecular Formula = Empirical Formula × 2

Molecular Formula = C_8H_{18}

Example 3: Cobalt (II) nitrate is a hydrate with a chemical formula of $Co(NO_3)_2 \bullet xH_2O$. When the 2.45 g of hydrate is heated, 1.54 g of residual is left behind. Determine the number of hydrate unit for cobalt (II) nitrate.

Mass of $Co(NO_3)_2$ = 1.54 g

Mass of H_2O released = 2.45 g − 1.54 g = 0.91 g

$$n = \frac{1.54\ \cancel{g}}{182.95\ \cancel{g}/mol} = 0.0084176004\ mol\ Co(NO_3)_2$$

$$n = \frac{0.91\ g}{18.02\ g/mol} = 0.0504994451\ mol\ H_2O$$

$$\frac{n_{H_2O}}{n_{Co(NO_3)_2}} = \frac{0.0504994451\ mol\ H_2O}{0.0084176004\ mol\ Co(NO_3)_2} \approx \frac{6\ mol\ H_2O}{1\ mol\ Co(NO_3)_2}$$

Molecular Formula = $Co(NO_3)_2 \bullet 6\ H_2O$

Assignment
3.5 & 3.6 pg. 87−88 #40 to 54 (even)

3.7: Chemical Reactions and Chemical Equations

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: (*For a new pure substance formed*)

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

Reactants: - chemicals that go 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.

Chemical Equation: - uses chemical symbols to represent what happens in a chemical reaction.

States of Chemicals: - (*s*) solid, (*l*) liquid, (*g*) gas, (*aq*) aqueous – dissolved in water

There are 5 basic types of chemical reactions:

1. **Formation, Composition or Synthesis**
   (Many Elements → Compound or Many Reactants → Single Product)

   Examples: $2\ Mg_{(s)} + O_{2\,(g)} \rightarrow 2\ MgO_{(s)}$ $SO_{2\,(g)} + H_2O_{(l)} \rightarrow H_2SO_{3\,(aq)}$

2. **Deformation or Decomposition**
   (Compound → Many Elements or Single Reactant → Many Products)

   Examples: $2\ Al_2O_{3\,(s)} \rightarrow 4\ Al_{(s)} + 3\ O_{2\,(g)}$ $H_2CO_{3\,(aq)} \rightarrow H_2O_{(l)} + CO_{2\,(g)}$

3. **Single Replacement** **(Element + Compound → Element + Compound)**

   Example: $2\ AgNO_{3\,(aq)} + Cu_{(s)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$

4. **Double Replacement** **(Compound + Compound → Compound + Compound)**

   Example: $AgNO_{3\,(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3\,(aq)}$

5. **Hydrocarbon Combustion** **(Hydrocarbon + Oxygen → Carbon Dioxide + Water)**

   Example: $CH_{4\,(g)} + 2\ O_{2\,(g)} \rightarrow CO_{2\,(g)} + 2\ H_2O_{(g)}$

Balancing Chemical Equations

Coefficient: - the number in front of the chemical formula that indicates the number of moles, atoms or molecules involve in a chemical reaction.
- the absence of number in front of chemical means the coefficient is 1.

Example: $2\ H_{2\,(g)} + 1\ O_{2\,(g)} \rightarrow 2\ H_2O_{(g)}$

To Predict Products and Balance Chemical Equations:

1. Write the correct chemical formulas for all products and reactants with proper subscripts. The presence of metals or ionic compounds indicates that we will need to use ions and charges to form any products.
2. For hydrocarbon combustion, balance in the order of C, H, and then O. The product, H_2O, is always in gaseous form unless otherwise stated. (It's usually quite hot in combustion.)
3. For other type of reactions, balance the equation for each type of cations and anions. Do NOT break up the polyatomic ions. Water may be written as HOH (H^+ and OH^-) in single and double replacements.
4. Check with the Solubility Table (see Section 4.2) and the Table of Elements for the states of chemicals.

Example 1: Predict the product(s) along with the states, indicate the type of reaction, and balance the following chemical reactions.

a. Sulfur trioxide gas is produced from its elements.

Formation: $S_{8\,(s)} + 12\ O_{2\,(g)} \rightarrow 8\ SO_{3\,(g)}$

b. A solid piece of zinc is immersed in an iron (III) chloride solution.

Single Replacement: $3\ Zn_{(s)} + 2\ FeCl_{3\,(aq)} \rightarrow 3\ ZnCl_{2\,(aq)} + 2\ Fe_{(s)}$

Zn^{2+} Fe^{3+} Cl^{-}

c. Propane ($C_3H_{8\,(g)}$) is burned in a gas barbecue.

Hydrocarbon Combustion: $C_3H_{8\,(g)} + 5\ O_{2\,(g)} \rightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{(g)}$

d. Chlorine gas is bubbled through a copper (II) iodide solution.

Single Replacement: $Cl_{2\,(g)} + CuI_{2\,(aq)} \rightarrow CuCl_{2\,(aq)} + I_{2\,(s)}$

Cl^{-} Cu^{2+} I^{-}

e. Ammonia gas is decomposed into its elements.

Decomposition: $2\ NH_{3\,(g)} \rightarrow N_{2\,(g)} + 3\ H_{2\,(g)}$

f. Sulfuric acid is neutralized by sodium hydroxide solution.

Double Replacement: $H_2SO_{4\,(aq)} + 2\ NaOH_{(aq)} \rightarrow 2\ HOH_{(aq)} + Na_2SO_{4\,(aq)}$

H^{+} SO_4^{2-} Na^{+} OH^{-}

g. Propanol ($C_3H_7OH_{(l)}$) is accidentally ignited.

Hydrocarbon Combustion: $C_3H_7OH_{(l)} + 9/2\ O_{2\,(g)} \rightarrow 3\ CO_{2\,(g)} + 4\ H_2O_{(g)}$
(Multiply Coefficients by 2) $2\ C_3H_7OH_{(l)} + 9\ O_{2\,(g)} \rightarrow 6\ CO_{2\,(g)} + 8\ H_2O_{(g)}$

h. Lead (II) nitrate solution is reacted with chromium (III) sulfate solution.

Double Replacement: $3\ Pb(NO_3)_{2\,(aq)} + Cr_2(SO_4)_{3\,(aq)} \rightarrow 2\ Cr(NO_3)_{3\,(aq)} + 3\ PbSO_{4\,(s)}$

Pb^{2+} NO_3^{-} Cr^{3+} SO_4^{2-}

i. Octane ($C_8H_{18\,(l)}$) is combusted in an automobile.

Hydrocarbon Combustion: $C_8H_{18\,(l)} + 25/2\ O_{2\,(g)} \rightarrow 8\ CO_{2\,(g)} + 9\ H_2O_{(g)}$
(Multiply Coefficients by 2) $2\ C_8H_{18\,(l)} + 25\ O_{2\,(g)} \rightarrow 16\ CO_{2\,(g)} + 18\ H_2O_{(g)}$

Assignment
3.7 pg. 88 #59, 60

3.8: Amounts of Reactants and Products

Stoichiometry: - the calculation of quantities of chemicals in a chemical reaction.

Stoichiometry Quantities: - when all quantities of the chemical are consumed.

Gravimetric Stoichiometry: - stoichiometry that involves quantities of masses.

Mole Ratio: - the ratio of coefficients between the required chemical and the given chemical.

Example 1: Interpret the chemical equation $4\ NH_{3\ (g)} + 7\ O_{2\ (g)} \rightarrow 4\ NO_{2\ (g)} + 6\ H_2O_{(g)}$ in terms of

a. moles. b. molecules. c. masses.

	$4\ NH_{3\ (g)}$	$7\ O_{2\ (g)}$	$4\ NO_{2\ (g)}$	$6\ H_2O_{(g)}$
a.	**4 moles of NH_3**	**7 moles of O_2**	**4 moles of NO_2**	**6 moles of H_2O**
b.	**4 molecules of NH_3**	**7 molecules of O_2**	**4 molecules of NO_2**	**6 molecules of H_2O**
c.	$m = nM$ $m = (4\ \cancel{\text{mol}})(17.034\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 68.136\ \text{g}}$	$m = nM$ $m = (7\ \cancel{\text{mol}})(32.00\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 224.0\ \text{g}}$	$m = nM$ $m = (4\ \cancel{\text{mol}})(46.01\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 184.0\ \text{g}}$	$m = nM$ $m = (6\ \cancel{\text{mol}})(18.02\ \text{g}/\cancel{\text{mol}})$ $\boldsymbol{m = 108.1\ \text{g}}$

Gravimetric Stoichiometry Procedure:

1. Predict the products and balance the chemical equation.
2. Put all the information given under the appropriate chemicals and determine the molar masses of the chemical involved.
3. Find the moles of the given chemical. $\left(n = \dfrac{m}{M}\right)$
4. Find the mole of the required chemical using mole ratio.
$$\left(\textbf{mol of require} = \textbf{mol of given} \times \frac{\textbf{require coefficient}}{\textbf{given coefficient}}\right)$$
5. Convert mole of the required chemical to its mass equivalence. **($m = nM$)**

Example 2: 4.35 mol of $PCl_{5\ (g)}$ is decomposed into its elements. Write a balance equation and determined the amount of chlorine produced.

$$\underline{4}\ PCl_{5\ (g)} \rightarrow P_{4\ (s)} + \underline{10}\ Cl_{2\ (g)}$$

4.35 mol ? mol

$$n_{Cl_2} = 4.35\ \cancel{\text{mol PCl}_5} \times \frac{10\ \text{mol Cl}_2}{4\ \cancel{\text{mol PCl}_5}} = 10.875\ \text{mol Cl}_2$$

$$\boldsymbol{n_{Cl_2} = 10.9\ \text{mol}}$$

Example 3: Determine the mass of carbon dioxide formed when 50.0 kg of butane ($C_4H_{10\,(g)}$) is burned.

$$2\,C_4H_{10\,(g)} + 13\,O_{2\,(g)} \rightarrow 8\,CO_{2\,(g)} + 10\,H_2O_{(g)}$$

C_4H_{10}: 50.0 kg, 58.12 g/mol; CO_2: ? g, M = 44.01 g/mol

① $n_{C_4H_{10}} = \dfrac{50.0\ \cancel{k}g}{58.12\ \cancel{g}/mol} = 0.8602890571$ kmol C_4H_{10}

② $n_{CO_2} = 0.8602890571\ \cancel{kmol\ C_4H_{10}} \times \dfrac{8\ mol\ CO_2}{2\ \cancel{mol\ C_4H_{10}}} = 3.441156228$ kmol CO_2

③ $m_{CO_2} = nM = (3.441156228\ k\cancel{mol}\ CO_2)(44.01\ g/\cancel{mol})$ $\quad$ $\boldsymbol{m_{CO_2} = 151\ kg}$

Example 4: Barium chloride solution was mixed with an excess sodium phosphate solution. What was the mass of barium chloride solid needed in the original solution to form 3.21 g of precipitate?

$$3\,BaCl_{2\,(aq)} + 2\,Na_3PO_{4\,(aq)} \rightarrow Ba_3(PO_4)_{2\,(s)} + 6\,NaCl_{(aq)}$$

$BaCl_2$: ? g, 208.23 g/mol; $Ba_3(PO_4)_2$: 3.21 g, M = 601.93 g/mol

① $n_{Ba_3(PO_4)_2} = \dfrac{3.21\ \cancel{g}}{601.93\ \cancel{g}/mol} = 0.005332846$ mol $Ba_3(PO_4)_2$

② $n_{BaCl_2} = 0.005332846\ \cancel{mol\ Ba_3(PO_4)_2} \times \dfrac{3\ mol\ BaCl_2}{1\ \cancel{mol\ Ba_3(PO_4)_2}} = 0.015998538$ mol $BaCl_2$

③ $m_{BaCl_2} = nM = (0.015998538\ \cancel{mol}\ BaCl_2)(208.23\ g/\cancel{mol})$ $\quad$ $\boldsymbol{m_{BaCl_2} = 3.33\ g}$

3.9: Limiting Reagents

Excess: - the reactant with more than enough amount for the reaction.

Limiting Reagent: - the reactant with the smaller amount (after taken account of the mole ratio) for the reaction.

Note: A limiting reagent question will always have enough information to find the moles of both reactants.

Example 1: A loaf of banana bread recipe calls for 3 eggs, 2 cups of flour, 1 cup of sugar and 3 bananas. How many loaves of banana bread can we make if we have 20 eggs, 11 cups of flour, 7 cups of sugar and 20 bananas? What is the limiting ingredient?

$$3\text{ Eggs} + 2\text{ Flour} + 1\text{ Sugar} + 3\text{ Bananas} \rightarrow 1\text{ Banana Bread}$$

20 Eggs, 11 Flour, 7 Sugar, 20 Bananas → ?

$\dfrac{20\ \cancel{Eggs}}{3\ \cancel{Eggs}/loave} = 6.67$ loaves $\qquad \dfrac{11\ Flour}{2\ Flour/loave} = $ **5.5 loaves** $\qquad \dfrac{7\ Sugar}{1\ Sugar/loave} = 7$ loaves $\qquad \dfrac{20\ Bananas}{3\ Bananas/loave} = 6.67$ loaves

(Minimum – Flour is the limiting ingredient)

Steps to deal with Limiting Reagent Problems:

1. Assume one of the reactants is the limiting reagent and determine its mole amount.
2. Determine the mole amount of the other reactant.
3. Use the mole amount of the assumed limiting reagent and the mole ratio; calculate the mole amount of the other reactant actually needed.
4. If the mole amount of the other reactant is smaller than what is needed, then our assumption was wrong. The other reactant is the limiting reagent.
5. If the mole amount of the other reactant is bigger than what is needed, then our assumption was correct. It means that the other reactant is the excess.

Example 2: 5.00 g of phosphorus is reacted with 15.00 g of chlorine gas to produce phosphorus trichloride. Determine the mass of the product produced.

$P_{4\,(s)}$	+	$6\ Cl_{2\,(g)}$	$\rightarrow$	$4\ PCl_{3\,(s)}$
5.00 g		15.00 g		? g
M = 123.88 g/mol		M = 70.90 g/mol		M = 137.32 g/mol

Since there is enough information to determine the moles of two reactants, we need to determine which one is the limiting reagent.

① $n_{P_4} = \dfrac{m}{M} = \dfrac{5.00\ \cancel{g}}{123.88\ \cancel{g}/mol} = 0.04036\ldots$ mol P_4 ② $n_{Cl_2} = \dfrac{m}{M} = \dfrac{15.00\ \cancel{g}}{70.90\ \cancel{g}/mol} = 0.2115\ldots$ mol Cl_2

Let's assume P_4 is the limiting reagent. Calculate the mol Cl_2 actually needed.

③ $n_{Cl_2} = 0.0403616403\ \cancel{mol\ P_4} \times \dfrac{6\ mol\ Cl_2}{1\ \cancel{mol}\ P_4} = 0.2421698418$ mol Cl_2 needed

<u>But we don't have 0.2421698418 mol of Cl_2,</u> <u>we only have 0.2115655853 mol of Cl_2</u>. Therefore, Cl_2 is the limiting reagent. (*Note: the limiting reagent is <u>NOT</u> always the chemical with the smaller number of moles. You have to always compare like we did above.*)

Now, we calculate the moles of PCl_3 formed by using moles of limiting reagent Cl_2.

④ $n_{PCl_3} = 0.2115655853\ \cancel{mol\ Cl_2} \times \dfrac{4\ mol\ PCl_3}{6\ \cancel{mol}\ \cancel{Cl_2}} = 0.1410437236$ mol PCl_3

Finally, we determine the mass of PCl_3 produced.

⑤ $m_{PCl_3} = nM = (0.\ 1410437236\ \cancel{mol}\ PCl_3)(137.32\ g/\cancel{mol}) = 19.3664852$ g $\quad$ **m_{PCl_3} = 19.4 g**

3.10: Reaction Yield

<u>Percent Yield</u>: - sometimes call reaction yield
- the proportion of the actual yield to the theoretical yield.

$$\boxed{\%\ \textbf{Yield} = \frac{\textbf{Actual}}{\textbf{Theoretical}} \times 100\%}$$

Example 1: If the actual yield of the last example was 18.2 g, what is the percent yield of the reaction?

$$\%\ \text{Yield} = \frac{\text{Actual}}{\text{Theoretical}} \times 100\%$$
$$= \frac{18.2\ g}{19.4\ g} \times 100\%$$

% Yield = 93.8%

<u>Assignment</u>
3.8 pg. 89−90 #64 to 78 (even)
3.9 pg. 90 #82, 84, 86
3.10 pg. 91 #90, 92