Chemistry Notes

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## Unit 1: Matter

**Matter:** Amount of substance (kg)
**Volume:** Amount of space (ml)
If the temperature of a substance is increased, then the volume will also increase because the particles vibrate and the space increases.
**Temperature:** Sum of kinetic energy in particles
**Density:** Amount of substance in a certain volume
Density=Mass/Volume
Density will decrease as the volume increases

**What is Kinetic Energy?**

- Definition: Kinetic energy is the energy possessed by an object or particles due to their motion.
- K.E → Movement of particles and temperature, so that we can create
- Total K.E of a particle → Temperature

- **K.E is divided into 2 types:**
  1. Vibrational Kinetic Energy: This is the energy due to the vibration of particles. For example, in a solid, atoms vibrate around fixed positions.
  2. Translational Kinetic Energy: This energy is due to the movement of particles from one place to another. For example, in gasses, molecules move freely and collide with each other.

- When temperature increases → Kinetic Energy (K.E) increases → Volume increases.

**- Temperature and Volume Relationship:**
- Hypothesis: **If** the temperature of a substance is increased, **then** the volume of that substance will also increase **because** as the particles gain more energy they start to vibrate more vigorously causing them to take up more space.
- As the particles vibrate more, they push against each other and occupy more space, leading to an increase in volume.

**Hypothesis**: IF, THEN, BECAUSE

**Measuring Mass**
For solids, you put the object on a balance scale.
For liquids, you put the container on balance, then reset the balance scale, so it is ‘tared’, then you pour the liquid and record the reading.

**Measuring Volume**
For liquids, you use a measuring cylinder, and if they have a lower meniscus, the volume will be measured from the lower side. The same goes for an upper meniscus

**Dalton's Atomic Theory**
1. All substances are made up of particles called atoms, the different states of matter can be explained by their extent of separation and kinetic energy.
2. All atoms of an element are identical and every atom of different elements has different properties.
3. When atoms combine chemically, they form compounds with unique properties.
4. Atoms combine in a fixed ratio when combining a compound.

**Mass, Energy and Temperature**

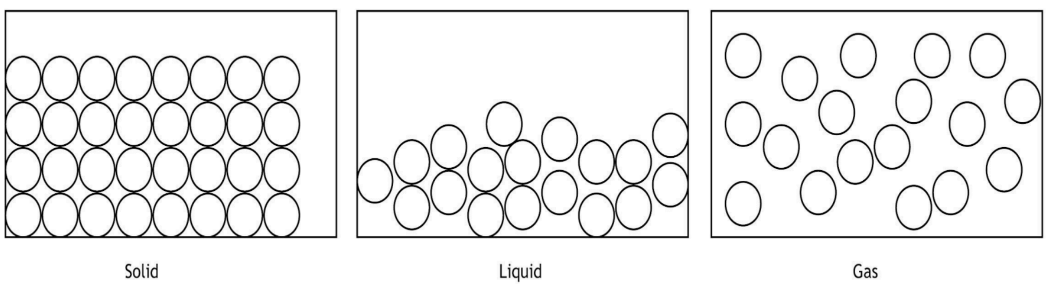
- When a substance is heated, its temperature rises which means the kinetic energy also rises. The rise of kinetic energy increases the vibrations and spaces between the particles which lead to change of state.

**Molecular Arrangements**

**Solids:** Compact/tightly packed, strong intermolecular bonds, fixed shape and fixed volume

**Liquids**: Loosely packed, moderate intermolecular bonds, do not have a fixed shape, fixed volume

**Gas:** Particles move around freely, weak bonding, do not have a fixed shape, do not have a fixed volume




**Solids To Gas**

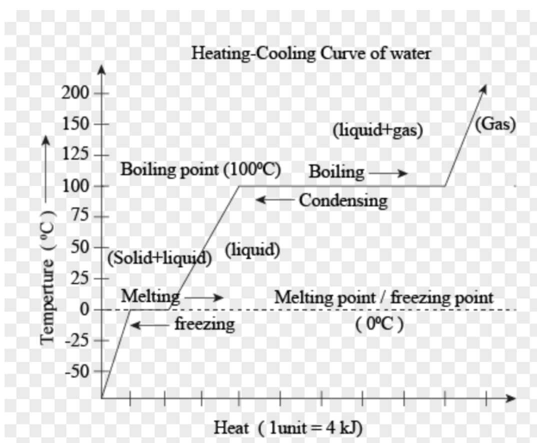
The particles in a solid are tightly packed so when the temperature is increased, the particles start to vibrate and move away from each other and their volume increases. This is when the solid starts to melt into a liquid, in which in the molecular arrangements the molecules are further apart. When the temperature is increased more, the particles start to move rapidly. This causes the particles to move further apart from each other. When the liquid is heated to a certain point, particles start to evaporate and become a gas.

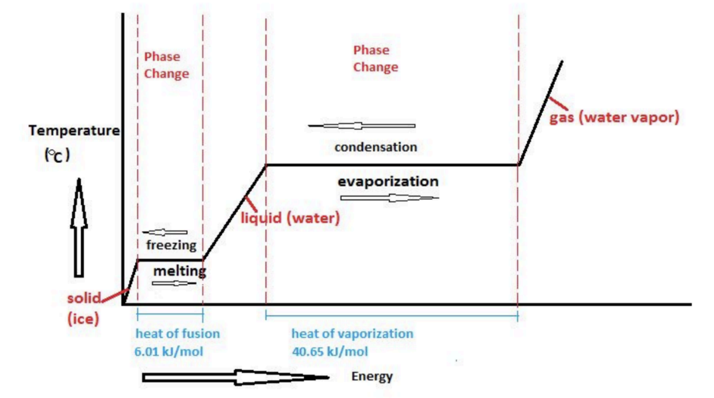
**Liquid to Solids**

The process of conversion of liquid to solid is called freezing. The liquid molecules are originally further apart, they have a fixed volume and slide against each other to move. When the temperature of a liquid is decreased, the energy absorbed will be used in order to form bonds between the particles. The kinetic energy of the liquid will be decreased and ultimately since bonds will be formed, the particles of a liquid will come close together and form a solid in which particles are close together, they vibrate and have a fixed volume.

## Heating Curve

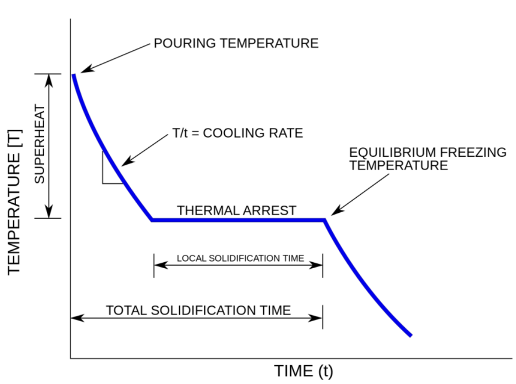
Energy is used to break bonds and change into liquid/gas

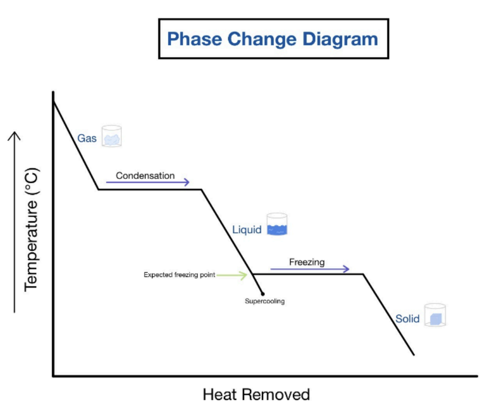



## Cooling Curve

Energy is used to strengthen bonds and change into liquid/solid



**Physical Change**

- It is a change that can be observed by our 5 senses
- There is no new substance formed
- Generally physical changes are reversible
- Eg: Melting, Boiling, Evaporation, Freezing

**Chemical Change**

- It can not be observed by our 5 senses
- New substance are formed
- Generally chemical changes are irreversible
- Eg. Hydrogen Sulphide

**Melting:** when a solid converts into a liquid.
**Boiling:** when a liquid converts into a gas.
Boiling happens to the entire liquid but evaporation only happens on the top layers, it is a slow process which slowly converts the top layers into gas until all is gas.
**Condensation:** when a gas converts into a liquid.
**Freezing:** liquid to a solid
**Sublimation:** solids to gasses directly.
**Deposition:** gasses convert directly into solids.

★ The IV lies on the x-axis.
★ The DV lies on the y-axis.
★ Write the units along with the quantities.
★ The title of the graph is the impact of the IV on the DV.

# Unit 2: Separation techniques

1. Types of mixtures: Homogeneous and heterogeneous mixtures
2. Solid - Solid separation
3. Solid - Liquid separation
4. Liquid - Liquid separation
5. Qualitative analysis: Chromatography

**Types Of Mixtures**
Homogenous: It is when the outcome is one type of phase (state of matter)
Heterogeneous: It is when the outcome is more than one phase (state of matter)

**Miscible:** liquids that mix. E.g Acid + water, alcohol + water and gasoline
**Immiscible:** liquids that cannot mix. E.g oil + water

**Solid-Solid Separation**
<u>Magnetic Separation:</u> Uses a magnet to separate magnetic materials from non-magnetic ones.

Pros: Quick, effective for magnetic materials
Cons:

1) Only works for mixtures with magnetic substances inside
2) If there are other non-magnetic substances they won't be separated
3) It won't work for liquid-liquid
4) Solid submerged in water would reduce magnetic strength – difficult
5) If there are 2 or more magnetic substances they won't be separated
6) Fine particles won't be separated out – too small to have any attraction

Safety Considerations

1) Metals can snap together fast, be careful

2) Make sure to wear a face masks and goggles to protect your eyes from small particles
3) Do not bring people on electronic support e.g pacemakers as the metal would disrupt it

<u>Sieving:</u> Uses a sieve to separate particles based on density.

Pros: Simple, cost effective for size based separation

Cons:

1) Ineffective for liquid-solid mixtures and liquid-liquid mixtures
2) Does not separate objects with same/similar densities
3) Small particles will be left
4) Hard to collect small particles after separation

**Solid-Liquid Separation**

<u>Filtration(heterogeneous mixtures)</u>

- Fold the filter paper into a semicircle, then fold again into a quarter-circle.
- Open one layer to form a cone.
- Place the filter cone inside a funnel using a bit of water to stick it to the edges firmly.
- Position the funnel on a tripod stand.
- Ensure the funnel tip aligns with the beaker placed below to collect the solvent.
- Pour the mixture into the filter cone.
- The solid remains on the filter paper, and the liquid drips into the beaker.

If a heterogeneous mixture is passed through a filter, then the solid will be separated from the liquid because the filter paper allows only the liquid particles to pass through while trapping the solid particles.

**Hypothesis:**

(homogenous)**If** we put salt water in the funnel cone, **then** the water and salt will not separate **because** the salt has dissolved in the water.

**Hypothesis:**
(heterogeneous) If we put sand water in the funnel cone ( filtration process) then the sand will separate from the water because the size of the sand particles are larger than the pores of the filter paper, so the water will pass and sand stays behind.

Advantages:
1) Simple and easy
2) Allows us to separate solid-liquid mixtures

Disadvantages:
1) Sometimes pores are clogged due to large particle size
2) Does not separate mixtures where the solute is dissolved
3) Does not separate liquid-liquid or solid-solid
4) Is slow for viscous liquids
5) Does not filter out bacteria

Safety Considerations:
1) Do not pour water too fast or it will spill
2) Wear goggles, gloves and coat so that if solvent falls your clothes won't get dirty
3) Keep the end of the funnel on the side of the beaker so that when the solvent falls it wont splash

**DATA CONSIDERATIONS: (OVERALL ASPECT)**
**VERY IMPORTANT**

Everything:

1) Take 3 **trials**, if not similar then 5, then taken the **mean**

2) If there is an evident **outlier**, do not take it into account

3) Make sure to have a good range of IV trials to see if the relationship is constant throughout

4) Keep IV **intervals** same to see exact correlation between DV and IV

5) Make a table -- **IV will always be on the left** column, **DV always on right column**

6) Keep **units on the top**, not next to each value

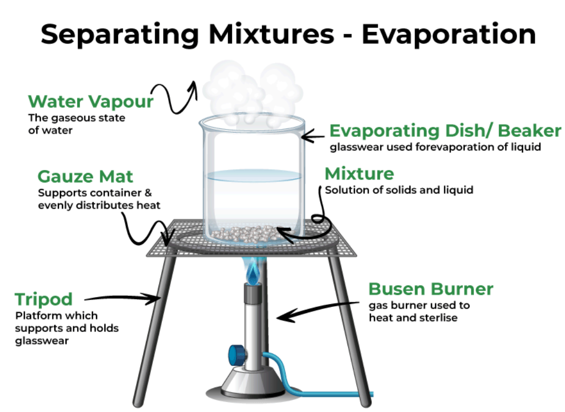
Graph:

1) Always keep intervals between each box same e.g first box is = 100 units, 2nd box should be the same

2) Y axis will have dependent variable and X axis with Independent

<u>Evaporating to drying(homogenous mixtures)</u>

- Place a tripod stand and position a wire gauze on it.
- Put the heating dish containing the liquid mixture on the wire gauze on the beaker.
- Use a bunsen burner under the tripod stand to heat the dish.
- A beaker can be placed nearby to collect condensed vapors if needed.
- The liquid evaporates, leaving the solid residue in the china dish.




(the liquid which is evaporated is lost)
If a liquid mixture is heated in a china dish, then the liquid will evaporate, leaving the solid behind because heat causes the liquid to change into vapor, separating it from the solid.

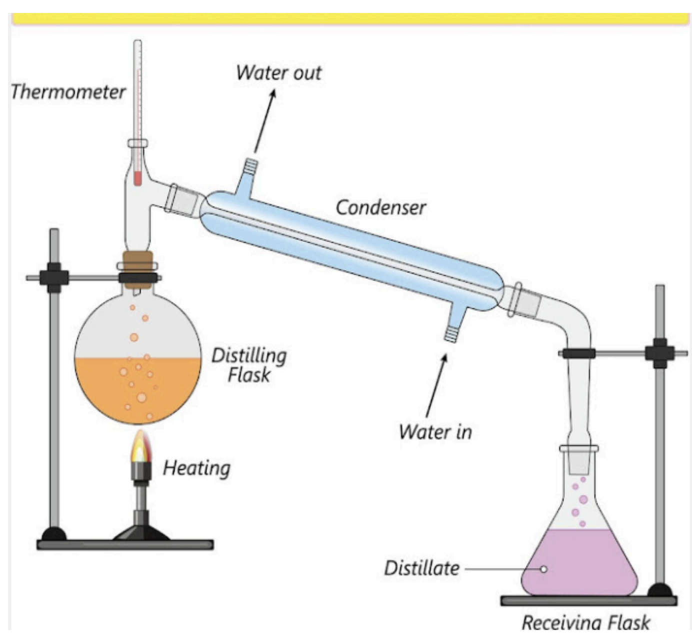
Advantages:
1) Can even separate dissolved solutes
2) Easily separates solid-liquid mixtures

Limitations
1) You lose the liquid
2) Solute may be degraded from the heat
3) Slow
4) Can be energy intensive if used a lot over time

Safety:
1) Keep flammable things away so they don't catch fire
2) Tie hair, wear coat, mask, gloves and goggles to protect yourself from vapours that may irritate your eyes and to keep your clothes safe from catching fire (clothe is flammable)

<u>Simple distillation (without losing the liquid)</u>

The mixture is separated due to the difference in temperature
- The mixture is placed in the round bottom flask where it is heated.
- When the boiling temperature of the liquid is reached, it evaporates and reaches the condenser
- The condenser cools the vapor and is separated out of the flask
- The solid will stay in the flask while the liquid evaporates and sits in the receiving/conical flask.



It works well for liquids with different boiling points and when we don't want to lose the liquids.
We call the liquid that has been separated the **distillate**.

Advantages:
1) Can separate liquid-liquid
2) We keep the liquid
3) Simple
4) We can even separate solid-liquid
5) Bacteria is killed (high heat melts the protein membrane)

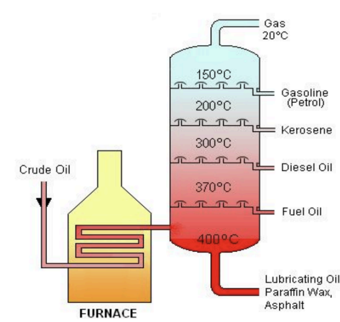
Limitations:
1) Can't separate solid solid
2) If they have too similar boiling points, closer than 25 degrees celsius, won't work.
3) Minerals are removed from the water
4) High boiling points can take a long time

Safety Considerations
1) Keep flammable things away
2) Tie hair, wear lab coat, goggles, gloves and face mask to protect from vapors and if your clothes get burnt

**Liquid-Liquid Separation**

<u>Fractional Distillation</u>

The mixture is separated due to difference in temperature.
The components whose boiling points come early, rise and are separated out and the ones which come later, will be separated out later.
Note: This technique is much more complex, so it's used for solutions which are more difficult to separate. Also, better for solutions with multiple liquids.



Advantages:

1) We can extract economically valued liquids such as petroleum and kerosene from crude oil
2) It allows us to separate a variety of different liquids at the same time.
3) The difference in boiling points can be less than 25 degree celsius and still work

Disadvantages:

1) Can't separate anything else
2) Energy intensive
3) Complex setup

Safety:

1) Keep flammable items away
2) Do not touch the system
3) Make sure to wear a lab coat, goggles, gloves, and mask (self-explanatory)

- **Miscible → Liquids that can mix**
- **Immiscible → Liquids that cannot mix**

<u>Separating Funnel</u>

Make the separating funnel system by holding it around 50 cm above the ground with a robert stand before putting a beaker underneath. Pour the immiscible mixture into the system from the top and the rest is yk.

Advantages:

1) Only method to separate immiscible liquids
2) Its quick and easy, can be done at home with the right equipment

Limitations:

1) Prone to human error

2) Can't be 100% that you time it right
3) Can't separate other mixtures

Safety:
1) Wear coat, goggles, mask so splashes won't hit you
2) Also do not keep the funnel too high because then the water will splash

**Qualitative Analysis**
<u>Chromatography</u>

Chroma - Colour
Graph - Representation

Representation of Colours
- Components of mixtures can be identified
- Chroma refers to color and graphy refers to representation
- Representation based on different colors

Every component has a fixed ratio which we call as Rf value;
Rf (retention factor) = Distance traveled by the component/Distance traveled by the solvent (solvent front)

It is always a number between 0 and 1, a smaller number means that the component has a greater affinity with the stationary phase (filter paper) and a greater number indicates that the component has a great affinity with the mobile phase (solvent).

**Equipment:**
- Chromatography Jar
- Filter paper
- Dropper

- Ink
- Glue
- Pencil
- Ruler
- Scissors
- Tape (optional)

**Steps:**

1. Draw a point on the filter paper around 10 cm away from the end.
2. Using the pencil, draw a line on the filter paper with a ruler.
3. Add the ink through the dropper and add the ink on the line.
4. Use the ruler to measure a distance of 5 cm from the bottom of the jar and place a small bit of tape there and mark it with the pencil.
5. Add water to the jar until it reaches the mark.
6. Place the paper through the hole till the edge of filter paper reaches the water (remember to stick the filter paper to one of the rods with glue so it stays in place).

**Pure substances:** They only have 1 type of atom/molecule/compound. Has fixed physical properties.

**Impure substances:** More than one type of atom/molecule/compound. It is when we add another substance to a pure substance and they mix **physically**. If they mix chemically they will form a new pure substance. Making a pure substance impure will alter its properties.

Solute → lesser quantity
Solvent → excess quantity
Solution → homogeneous mixture ( solute + solvent )

**Commercial Mixtures**
**Emulsions:** a mixture of two or more immiscible solvents stabilized by an emulsifying agent.

2 types:
**Oil in water** - Less oily (water is the solvent) e.g sauces, dressings, ointments
**Water in Oil** - more oily (oil is the solvent) e.g butter, creams
- We use an emulsifying agent to mix the 2 immiscible liquids together (usually water and oil)

**Colloids:** Mixture between two or more substances with a particle size of 1-1000 nanometers. E.g Milk, Jelly, fog, ink.

**Suspension:**
Mixture of substances that has a particle size **greater than 1000 nm**.
**Suspending agents** are used for stabilizing suspensions.
Particles more dense → gravity, settle down at the bottom → pulpy

## Unit 3: Mapping Matter

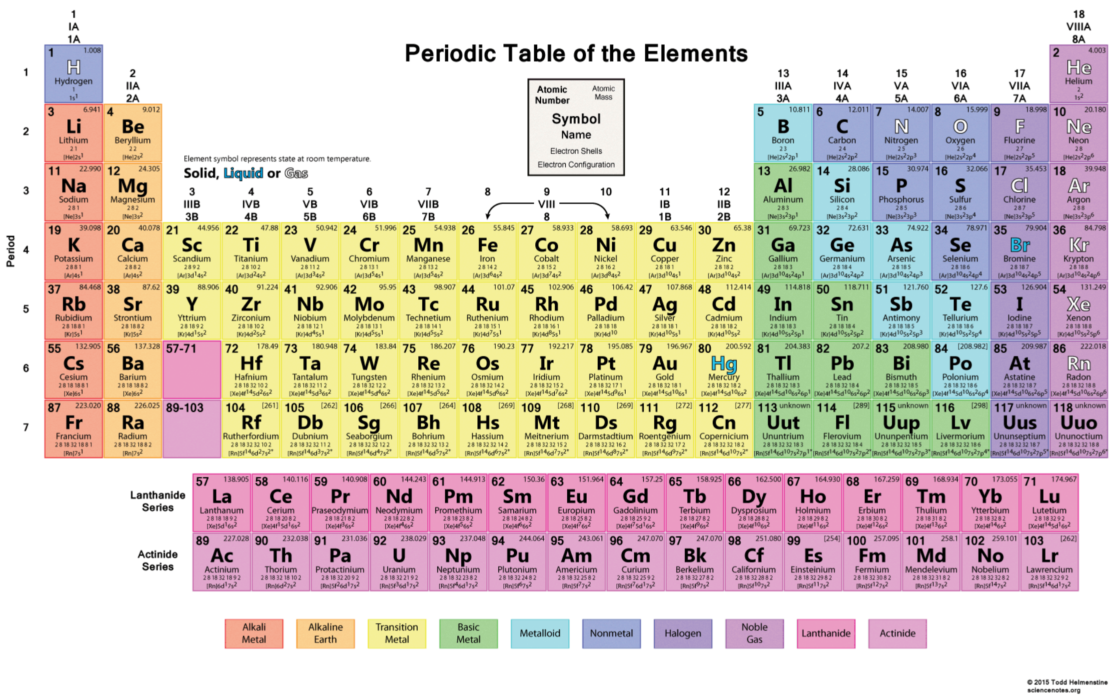
Matter exists in 3 different forms: Elements, Compounds, Mixtures/Solutions
Element: Substance that has the same type of atoms

- Monatomic: Elements that exists as a single atoms, eg. Helium(He), Neon (Ne)
- Diatomic: Elements that exist as two atoms, eg. Hydrogen($H_2$), Oxygen ($O_2$)
- Triatomic: Elements that exist in forms of three atoms, eg.Ozone($O_3$)
- Polyatomic: Elements that exist in forms of change, eg. Diamonds, Graphite

**Periodic Table**
The capital letters tells us the amount of atoms in an element
The atomic number of an element is its identity

The atomic mass can change

- The atomic number is the number of protons
- The atomic mass is the number of protons and neutrons
- The electrons are the same as the number protons (only if there is no charge)




Octet Rule: Every element focuses on completing their outermost shell with 8 electrons. The first shell will always need 2 electrons to complete so elements which only have one shell like hydrogen, only need 2 electrons to complete their outermost shell.

Groups: vertical representation of elements

Periods: horizontal representation of elements

Group numbers represent the number of electrons in the outermost shell, also known as valence electrons.

E.g Mg (**group 2**) → **2 valence electrons** in the outermost shell
E.g Helium (**group 8**) → **8 valence electrons** in the outermost shell

Group numbers tell us how many valence electrons there are and the valency.
Period number represents the number of shells.
E.g Sodium (period 3) 3 shells
There can be a maximum of two electrons in the first shell and eight in the rest (max 8 shells).

Names of Groups

1) Group 1: Alkali Metals - Highly reactive
2) Group 2: Alkali Earth Metals
3) Groups 3 - 6 are non metals
4) Group 7: Halogens - Highly reactive
5) Group 8: Noble Gases - Unreactive

Transition metals do not count as a group.
Metals form coloured compounds (they have a colour).

Properties Of Metals

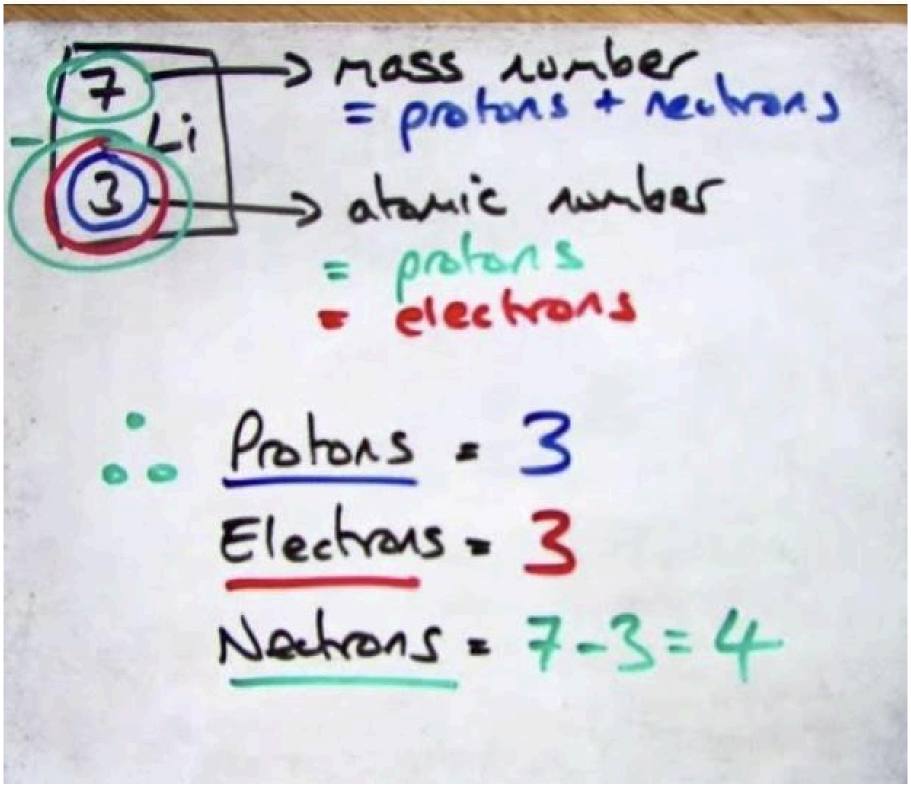
1) **Hard** - it can resist scratching, not easily bent broken or pierced
2) **Ductile** - Can be made into thin wires
3) **Malleable** - Can be hammered into different shapes without breaking
4) **Lustrous** - They are shiny
5) **Sonorous** - They make a sound when struck
6) **High Melting and Boiling Points** - self-explanatory
7) **Electric Conductivity** - Electricity can pass through it easily
8) **Heat Conductivity** - Heat passes and spreads through it fast due to a free electron which helps speed up the process
9) **High Density** - Very compact, this contributes to its hardness
10) **Magnetic Attraction** (sometimes) - some metals attract other metallic objects towards them

Features Of Modern Periodic Table

- Elements are arranged in increasing atomic number
- If you go left to right number of protons increase
- If you go up to down number of shells increase
- Divided into metals, nonmetals and metalloids
- Groups and Periods
- Properties show periodic trends
- Same group = similar properties

Valencies
Group 1→ +1
Group 2 → +2
Group 3 → +3
Group 4 → +/- 4
Group 5 → -3
Group 6 → -2
Group 7 → -1
Group 8 → 0

For transition metals:
ic→ bigger valency
ous→ smaller valency


7
Li
3
mass number
= protons + neutrons
atomic number
= protons
= electrons
Protons = 3
Electrons = 3
Neutrons = 7-3=4

| Property | Definition | Period Trend | Reason (Across a Period) | Group Trend | Reason (Down a Group) |
|---|---|---|---|---|---|
| **Atomic Radius** | Distance from the center of an atom's nucleus to the outermost electron | Decreases | More protons increase nuclear pull, pulling electrons closer | Increases | Addition of electron shells makes the atom larger |
| **Ionization Energy** | The energy required to remove an electron from an atom to form a positive ion | Increases (1-3), Decreases (5-7) | Protons increase nuclear pull, requiring more energy to remove electrons | Decreases | Outer electrons are farther from the nucleus, making them easier to remove |
| **Electronegativity** | The ability of an atom to attract electrons when it is part of a compound | Increases | More protons create a stronger attraction for bonding electrons | Decreases | Outer electrons are farther from the nucleus, reducing the ability to attract electrons |

**Trends**

Group: As the groups increase 1-2-3-4, the number of electrons in the outermost shells increases.

Periods: As the period increases 1-2-3-4, the number of shells increases.

Alkali - OH present

Acid - $H^+$ present

**Oxides:**

- Oxides are compounds made when an element reacts with oxygen. For example, when iron reacts with oxygen, it forms iron oxide (rust).
- **Metal oxides** tend to be **basic** (like sodium oxide, which can react with water to form a base).

- **Non-metal oxides** are **acidic** (like carbon dioxide, which can form an acid when dissolved in water).

2 types of oxides

1) Basic → Alkaline (has OH)
2) Acid → (Has $H^+$ ion)

Alkaline oxides are only made with metals, so groups 1 - 3.

We first make a reaction with Oxygen, then add $H_2O$

After we have made an oxide we have to make a compound with water to complete the process.

**For Groups 1 - 3**

Example: (tip: when groups 1 to 3, just add hydroxide at the end and then hydroxide is $OH^-$ so you can do the formula)

$Na^{+2} + O^{-2} \rightarrow NaO$

Na + O → Sodium Oxide

$NaO + H_2O$ → Sodium **Hydroxide**

$NaO + H_2O \rightarrow Na\mathbf{OH}$

**(Metal + O) + $H_2O$ → (Metal + OH)**

This is **alkaline**
Groups 4 - 7
They are acids, all of them will undergo the same process but an acid will be made in the end not a hydroxide. Group 7 is the only exception, it only has to make a compound with Hydrogen directly to make an acid.

Group 4 - 6 example:

Nitrogen + Oxygen → Nitrogen Dioxide
$N_2$ + $O_2$ → $NO_2$
After that add $H_2O$

**Nitrogen** Dioxide + Water → **Nitric** Acid
$NO_2$ + $H_2O$ → $HNO_3$ (we got this because nitric acid is a commercial name, which is = $HNO_3$)

Carbon + Oxygen → Carbon Dioxide
C + $O_2$ → $CO_2$

$CO_2$ + $H_2O$ → $H_2CO_3$
**Carbon** Dioxide + Water → **Carbonic** Acid

Group 7 Example
**Fluorine** + Hydrogen → Hydro**fluoric** Acid
$F_2$ + $H_2$ → HF

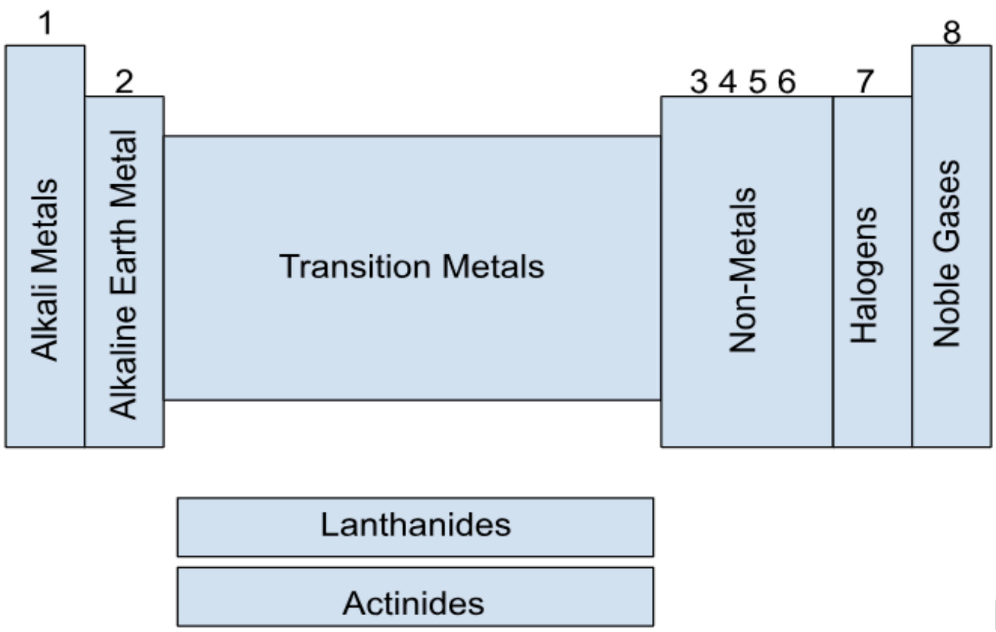
**Chlorine** + Hydrogen → Hydro**chloric** Acid (commercial name)
$Cl_2$ + $H_2$ → HCl

**In summary:**

- **Atomic radius**: Distance between nucleus and valence electron
- **Ionization energy**: Energy needed to react/ionize

First ionization energy is the energy required to remove the first electron, second ionization energy is the energy required to remove the second electron, and third ionization energy is the energy required to remove the third electron. When more electrons are taken, the protons can focus their attraction energy on a smaller number of electrons, so it's harder to take the rest.

- **Electronegativity**: Ability to attract electrons
- **Oxides**: Compounds formed when elements react with oxygen (they can be acidic or basic depending on the element).




Types of Reactions

1. **Addition reactions:**

The reaction in which the elements or compounds add up to give one product is called addition reaction

$A + B \rightarrow C$

- Magnesium + oxygen → Magnesium oxide (word equation)

  $Mg^{+2}\ O_2^{-2}$

  $Mg + O_2 \rightarrow MgO$

- Calcium Oxide + Carbon Dioxide → Calcium Carbonate

  $Ca^{+2}\ O^{-2} + CO_2 \rightarrow Ca^{+2}\ CO_3^{-2}$

  $CaO + CO_2 \rightarrow CaCO_3$

- Aluminum + Oxygen → Aluminum Oxide

  $Al^{+3} + O_2^{-2} \rightarrow Al_2O_3$

2. **Decomposition reaction:**

The type of reaction in which compounds break down into its respective elements (opposite of addition reaction). This reaction is usually caused by heat;

Calcium Carbonate —heat→ Calcium Oxide + Carbon Dioxide

$Ca^{+2}\ CO_3^{-2} \rightarrow Ca^{+2}\ O^{-2} + CO_2$

$CaCO_3$ —heat→ $CaO + CO_2$

3. **Combustion reaction: (burning)**

The type of reaction in which an organic compound (inflammable carbon based compound) catches fire in the presence of oxygen to release carbon dioxide and water;

Example carbon based compound: $CH_4$ (methane)

Methane + Oxygen —heat→ Carbon dioxide + Water + Heat

$CH_4 + O_2 \rightarrow CO_2 + H_2O$ + Heat

There are instances in which the supply of oxygen is limited, this causes incomplete combustion and instead of carbon dioxide, carbon monoxide is produced which is toxic;

Methane + Oxygen (limited) —heat→ Carbon Monoxide + water + heat

$CH_4 + O_2$ (limited) —heat→ $CO + H_2O$ + Heat

4. **Displacement reaction:**

   a. **Single Displacement Reaction**: The type of reaction in which one element displaces another one;

$A + BC \rightarrow AB + C$

Magnesium + hydrogen chloride → Magnesium chloride + Hydrogen

$Mg^{+2} + H^{+1}\ Cl^{-1} \rightarrow Mg^{+2}\ Cl^{-1} + H_2$

$Mg + HCl \rightarrow MgCl_2 + H_2$

Zinc (2) + Hydrogen sulphate → Zinc sulphate + hydrogen gas

$Zn^{+2} + H^{+1}\ SO_4^{-2} \rightarrow Zn^{+2}\ SO_4^{-2} + H_2$

$Zn + H_2SO_4 \rightarrow ZnSO_4 + H_2$

b) **Double displacement reaction:** The type of reaction in which all elements are displaced to form new compounds is called double displacement reaction;

The reaction in which all the elements are displaced.

$AB + CD \rightarrow AD + BC$

Sodium Hydroxide + Hydrochloric acid → Sodium chloride + Water

$Na^+\ OH^- + H^{+1}Cl^{-1} \rightarrow Na^+\ Cl^- + H_2O$

$NaOH + HCl \rightarrow NaCl + H_2O$

5. **Neutralization reaction:**
   The type of chemical reaction in which an alkali reacts with acid to form salt and water;

   **Sodium** Hydroxide + Hydro**chloric** Acid → **Sodium chloride** (salt) + Water

   $NaOH + HCl \rightarrow NaCl + H_2O$

# Term 2

## Unit 4 - Bonding

### Main Unit Topics

1) Intramolecular Bonding

a) Electronic Configuration
b) Covalent Bonding
c) Ionic Bonding
d) Coordinate Covalent Bonds
e) Lewis & Dot and Cross Structures
f) Metallic Bonding
g) Properties of Metals + Explanations

2) Intermolecular Bonding

a) Partial Charges
b) Dipoles
c) Dipole-dipole Interactions
d) Hydrogen Bonding

### Intermolecular Bonding

Intermolecular bonding is the bonding that occurs between the atoms of elements. There are 4 ways in which this can happen, **Covalent bonding, Ionic Bonding**, **Coordinate Covalent Bonding** and **Metallic Bonding.**

Electronic Configuration

You write the number of electrons for the first 4 shells in this way
K → Shell 1 (maximum capacity is 2)
L → Shell 2
M → Shell 3
N → Shell 4

**Examples:**

Q1) State the Electronic Configuration of Nitrogen
Nitrogen has 7 Electrons
K = 2
L = 5

Q2) State the Electronic Configuration of $Na^+$
Na (sodium) normally has 11 electrons. However, it is positively charged, meaning that 1 electron has been taken away. Thus, instead of there being 11 electrons there will be 10.
K = 2
L = 8
**I**

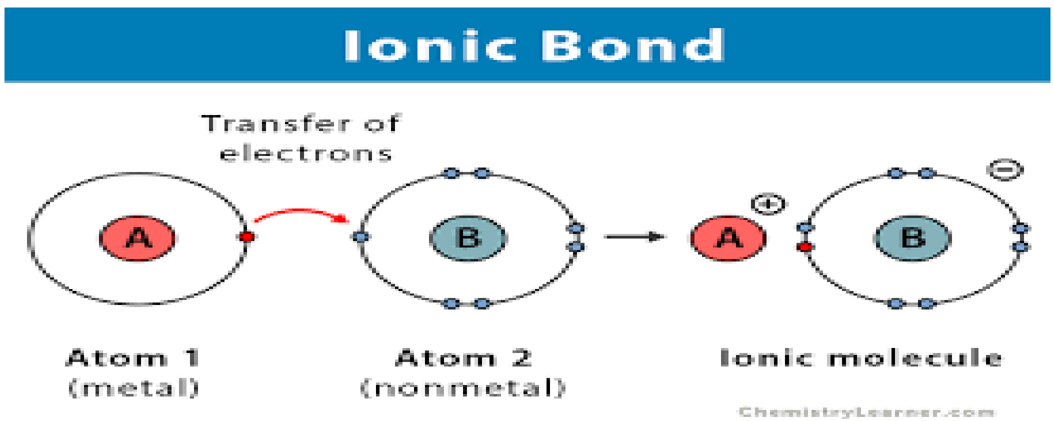
## Ionic Bonding

This type of bonding **can only occur between a metal and non-metal**, as opposite sign charges are required.

This is because in this type of bonding **one atom donates electrons to another to complete their own, and the other atom's octet.**

As a result, the **donor atom gains a positive charge** as it now has more protons than electrons (initially they were equal, now some electrons have gone away). The **receiver atom gains a negative charge** as it now has more electrons than protons.

Extra ‘lil fact:
**Ionic bonds** are **stronger than covalent bonds** as the **forces of attraction between the oppositely charged atoms** are very strong.

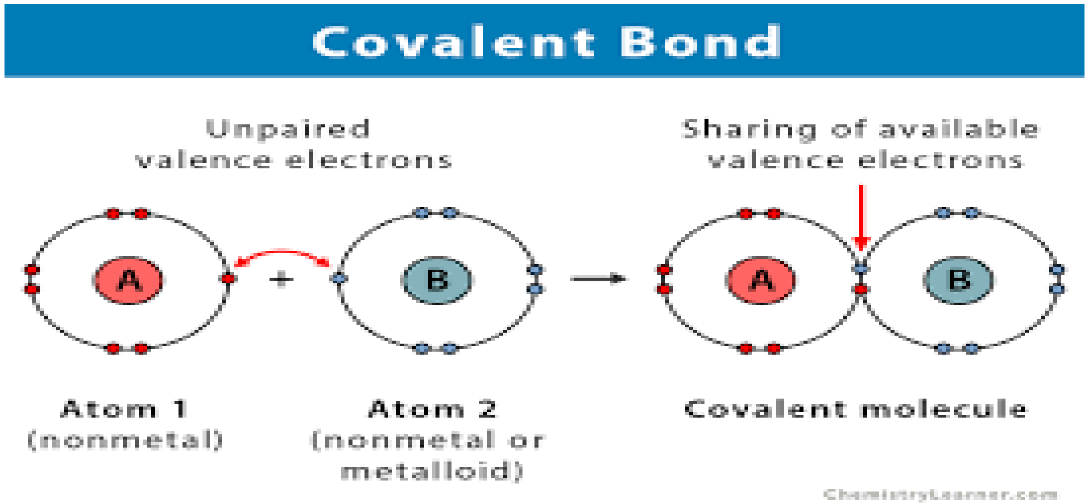



## Covalent Bonding

This type of bonding can **only occur between nonmetals.** This is because **electrons are shared** to complete the octet, and nonmetals can have a maximum of 3 valence electrons, so they wouldn’t be able to share to complete their octet (**max share is 3 electrons**).

The reason why 2 metals can't have a covalent bond is because in a covalent bond the atoms gain electrons, but metals want to lose electrons. Just saying.

There are 3 types of covalent bonds, **single covalent bond, double covalent bond,** and **triple covalent bond**. A single covalent bond has 1 electron from each atom being shared, a double bond has 2 electrons from each atom being shared and a triple bond has 3 electrons from each atom being shared.




## Dot and Cross & Lewis Structures

These structures are different ways of drawing the bonding between atoms.

The main difference is that in the **dot and cross** structures we draw **the entire atom,** including all of the inner electrons and shells. On the other hand, in the **Lewis structure** we **only draw the outer shell**.

Also, in the **dot and cross** structure while representing covalent bonds we **draw the electrons next to each other** (as shown in the above diagram), but the **Lewis structure** only shows a **line** (2 lines for double, and 3 lines for triple bonds).

## Coordinate Covalent Bonding

There are 2 prerequisites for this bonding. A **covalent bond** should be present, and that molecule should have **at least 1 lone pair**.

Lone Pair:
A pair of electrons that do not take part in the bonding.

If the molecule meets these conditions, then they can **donate that lone pair to an atom which needs it**, such as a $H^+$ ion. In a structure, this is represented as an **arrow** sprouting from the donor towards the receiver.

## Metallic Bonding

Metallic bonding is bonding between metals. You must first understand how this bonding takes place.

One of the unique properties of metal atoms is that they have **free electrons**. In this bonding, the **free electrons separate** from the atom, making a stream of free electrons and many positive charged atoms.

**Opposite charges → attracted together → create layers** of positive and negative charges.

**Electrons small → need less space → more atoms in a small space → strong bonding**
Opposite charge attraction is very strong, making this bonding even stronger.

## Properties of Metals

1) **High Density**

Since metals have very strong bonding, the atoms and electrons come closer to each other thus increasing the density

2) **Sonorus**

Sonorous means that when struck, metals make a ringing sound. This is because sound is generated through vibrations of the atoms across the whole element. Since the atoms are close, vibrations happen faster so it makes a sound.

3) **Malleable and Ductile**

This property allows metals, when struck, to change their shape and be rolled into thin sheets. When we strike the metal, the layer of alternating electrons and positive ions are moved up a bit, causing the same charges to be next to each other. These repel, causing that layer to break off and making the metal change shape.

4) **Heat and Electrical Conductivity**

Heat conductivity allows heat to be transferred throughout the metal faster, and electrical conductivity allows electrical charge to pass through the metal. The reason for the electrical conductivity property is because of the free electrons of the metal, which allow this to happen. Free electrons also allow heat conductivity, but another factor is that since the atoms are so close to each other they touch more so thermal energy is transferred quicker.

5) **High Melting and Boiling Points**

These are simply due to the fact that metals have strong intermolecular bonds due to the opposite charge's forces of attraction, thus making the bonds harder to break so they require more thermal energy.

6) **Lustrous**

The lustrous properties of metals make them shine when exposed to light. This is because the free electrons in the metal absorb light and reflect it back, thus making the metal shine.

7) **Hard**

Metals are strong and hard, which makes them harder to scratch, pierce and break. This is due to their high density.

## Intermolecular Bonding

This bonding is the bonding between the molecules themselves, not individual atoms. There are 2 types, **Dipole-Dipole Interactions** and **Hydrogen bonding.**

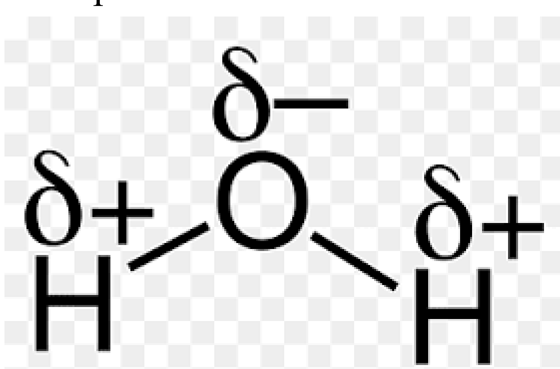
## Dipoles

A dipole is a molecule with **partial charges**.

Partial charges are only present in **covalent molecules (**molecules with covalent bonds).
When electrons are being shared, one atom will have **more protons** than the other, so their **electronegativity will be more**.
As a result, they will **pull harder on the electrons** so the shared **electrons go more towards that atom,** giving it a **partial negative charge**.

On the other hand, the **other atom will gain a partial positive charge.** If a molecule's atoms have partial charges, it is called a **dipole**.

Example




**Note: "$\delta$" is the sign for partial charge**

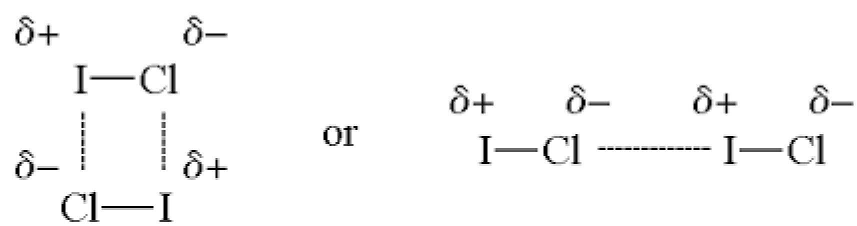
**Oxygen has 8 protons** whereas **Hydrogen has 1 proton**; This difference of protons causes the electrons to move more towards oxygen hence developing **partial negative charge on oxygen** and **partial positive charge on hydrogen.**

If both atoms are the same, they will pull the electrons with equal strength so no partial charges will be present, e.g $O_2$.

**The greater the difference between the protons** between the two elements, the **stronger the partial charges.**

## Dipole-Dipole Interactions

Since there are **many molecules** of the same compound, dipoles will **arrange** themselves according to the charges on themselves and the neighbouring molecule. This causes a **development of an attractive force** or a bond called **dipole-dipole** interactions;




As we can see, the **opposite partial charges are next to each other.**

## Hydrogen Bonding

**Hydrogen has the least amount of protons** in the entire periodic table, with only 1. This means that it has **extremely low electronegativity**. As a result, when it **forms a covalent bond with an extremely electronegative** element such as Oxygen, Chlorine, Fluorine or Nitrogen, the latter element will have a much greater pull on the electrons. This creates **very strong partial charges**, which are called hydrogen bonds.

Hydrogen bond is a bond where **hydrogen makes a bond with a very electronegative element.**

# Unit 5 – Chemicals

## Main Unit Topics

1) Theories of Acid Detection

a) Arrhenius
b) Lewis
c) Bronsted Lowry

2) Properties of Acids and Alkalis

a) Physical Properties
b) Chemical Properties (predicting chemical outcomes)

3) Detection Techniques

a) Detection of Hydrogen
b) Detection of Carbon Dioxide
c) Detection of Oxygen

4) Balancing Chemical Equations

5) Redox Reactions

a) Reduction
b) Oxidation

# Theories of Acid Detection

## 1) Arrhenius Theory

When we put a chemical inside an aqueous solution (water is involved) it **breaks down into its respective ions**. This is called a **dissociation reaction**. If the chemical releases Hydrogen Ions, it is an acid, and if it releases OH ions, it is an alkali. This is **reversible**.

Release $H^+$ → Acid

Release $OH^-$ → Alkali

Limitations to Arrhenius Theory

1. It is **limited to aqueous solutions** → we need an aqueous solution every time
2. It **doesn't explain the behavior of Ammonia** ($NH_3$). Although it releases Hydrogen it has the properties of an alkali.

## 2) Lewis Theory

The Lewis theory states that when forming chemical compounds, the compound/element which **accepts electrons is a Lewis Acid** and the compound which **donates electrons is a Lewis Base**.

Accepts Electrons → Lewis Acid
Donates Electrons → Lewis Base

Example
$NH_3 + H^+ \rightarrow NH_4^+$

**Ammonia donates** electrons to hydrogen to create a coordinate covalent bond in this reaction. Since ammonia is donating, it is a Lewis Base, and since **hydrogen is accepting** it is an acid.

Ammonia donates electrons → Lewis Base
Hydrogen accepts the electrons → Lewis Acid

3. Brønsted Lowry Theory

This theory states that when forming compounds, the compound which **donates a Hydrogen is a Bronsted Lowry Acid** and the compound which **accepts the Hydrogen is a Bronsted Lowry Base.**

**Donates** Hydrogen → Bronsted Lowry **Acid**
**Accepts** Hydrogen → Bronsted Lowry **Base**

Example
$NH_3 + HCl \rightarrow NH_4Cl$

In this reaction, hydrochloric acid (HCl) donated hydrogen to ammonia ($NH_3$). Since **Hydrochloric Acid donates a hydrogen** it is an acid, and since **ammonia accepts the hydrogen** it is an alkali.

Summary of Theories

| | **Acid** | **Base** |
|---|---|---|
| **Arrhenius** | Releases H+ in water | Releases OH- in water |
| **Bronsted-Lowry** | Donates Hydrogen | Accepts Hydrogen |
| **Lewis** | Accepts Electrons | Donates Electron |

## Properties of Acids and Alkalis

### Physical Properties

Litmus paper can identify PH level (color will tell us the strength).

Acids

1) Sour
2) Corrosive
3) Conduct Electricity
4) PH Level less than 7 (Less PH = Strong Acid)
5) Red Litmus

Things to know:

1) If an **acid is very strong** its **fumes alone can change the color of the litmus**
2) If we have to **submerge the litmus paper more** into the acid to get a reading, it is **weaker**.

Alkalis

1) Bitter
2) Corrosive
3) Conduct Electricity
4) PH Level more than 7 (More PH = Strong Alkali)
5) Blue Litmus

## Chemical Properties (only acid)

### 1) Reaction with Metals

Metal + Acid → Salt + Hydrogen
We can write it in another form to remember, let's imagine the metal is “x” and the chemical with hydrogen making an acid is “y”.
$x + Hy \rightarrow xy + H_2$

Examples:
$\mathbf{Mg} + H\mathbf{Cl} \rightarrow \mathbf{MgCl_2} + H_2$

$\mathbf{Na} + H_2\mathbf{SO_4} \rightarrow \mathbf{Na_2SO_4} + H_2$

The presence of hydrogen will always stay the same, but the **type of salt changes depending on the acid and metal.**

### 2) Reaction with Carbonates

Carbonate + Acid → Salt + Water + Carbon Dioxide
We can make a standard form by imagining the element with carbonate to be “x” and the compound making a bond with hydrogen as “y”.
$xCO_3 + Hy \rightarrow xy + H_2O + CO_2$

Examples:
$\mathbf{Mg}CO_3 + HCl \rightarrow \mathbf{Mg}Cl_2 + H_2O + CO_2$

$\mathbf{Al}_2(CO_3)_3 + H_2SO_4 \rightarrow \mathbf{Al}_2(SO_4)_3 + H_2O + CO_2$

In this type of reaction, the presence of water and carbon dioxide stay the same, the only thing that changes is the salt. **The element which is making a bond with carbonate makes a bond with the compound making an acid with hydrogen to make the salt.**

3) Neutralization Reaction

Acid + Alkali → Salt and Water

Examples
$H\mathbf{Cl} + \mathbf{Na}OH \rightarrow \mathbf{NaCl} + H_2O$

$H\mathbf{NO_3} + \mathbf{Al}(OH)_3 \rightarrow \mathbf{Al(NO_3)_3} + H_2O$

## Detection Techniques

### Detecting Hydrogen

We can detect hydrogen by taking a **burnt splint** (piece of paper on fire) and **holding it above the reaction.** If **hydrogen is present**, then there will be a **popping sound.**

### Detecting Carbon Dioxide

2 Ways:

1) **Burnt splint** → hold it over → carbon dioxide present → **flame extinguished**

2) **Limewater** → hold it over → carbon dioxide present → **milky/cloudy texture**.

*Lime Water* + Carbon Dioxide → **Milky/cloudy texture**
*Ca(OH)₂* + CO₂ → **CaCO₃** + H₂O + CO₂

This happens because the carbon dioxide comes into contact with the limewater.

## Detecting Oxygen

Burnt splint → oxygen present → **flame increase**

# Balancing Chemical Equations

Law of Conservation of Mass:
Matter cannot be created or destroyed it just changes form

When reacting with each other, atoms cannot magically disappear or appear out of nowhere. The number of atoms of each element has to be the same on both sides of the equation.

Example

Unbalanced Chemical Equation
$H_2 + O_2 \rightarrow H_2O$
On the left side there are 2 hydrogens and 2 oxygen, on the right there are 2 hydrogens and 1 oxygen, so we need to **balance it by multiplying the compounds.**

Balanced Chemical Equation

$2H_2 + O_2 \rightarrow 2H_2O$

When forming chemical equation we put **symbols next to an element to denote their state of matter:**

1) **Solid → (s)**
2) **Liquid → (l)**
3) **Gas → (g)**
4) **Aqueous → (aq)**

**Aqueous include salts**, and things that can **dissolve in water**

## Reduction and Oxidation

This concept determines whether an element/compound is **reduced** (undergone a **reduction**) or **oxidized** (undergone an **oxidation**).
There are **4 factors** for determining this:

| Reduction | Oxidation |
|---|---|
| **1) Loses an Oxygen** | **1) Gains an oxygen** |

Example;
$Na_2CO_3 + HCl \rightarrow NaCl + CO_2 + H_2O$
Carbon and Sodium lost oxygen, so they underwent reduction. Hydrogen gained oxygen and was oxidized.

| | |
|---|---|
| **2) Gains a Hydrogen** | **2) Loses a Hydrogen** |

Example;
$N_2 + H_2 \rightarrow NH_3$
Nitrogen gained hydrogen and so it was reduced.

| | |
|---|---|
| **3) Gains Electrons** | **3) Loses electrons** |

Example;

Na + Cl → NaCl

Sodium gave an electron to chlorine so it was oxidized. Chlorine took an electron so it was reduced.

4) **Oxidation state decreases** 4) **Oxidation state increases**

What is the oxidation state?

It is the **charge on the element/compound**.

Things to know:

1) All **solids** by themselves have an oxidation state (charge) of **0**
2) All **diatomic compounds** have an oxidation state of **0**, since they share electrons with each other
3) After making a bond, if an element **loses electrons** then the **oxidation state increases** (the charge on the element becomes positive) and if they **gain electrons** then the **oxidation state decreases** (the charge on the elements becomes negative)
4) Covalent compounds can exist as liquids or gases, and their sharing is temporary and unequal

$K + Cl_2 \rightarrow KCl$

Potassium previously had an oxidation state of 0, but after **donating an electron to Chlorine** its oxidation state **increased to 1,** so it was oxidized. **Chlorine gained an electron**, so its oxidation state went from 0 to -1, meaning it was **reduced**.

# Unit 6 – What Determines Chemical Change?

## Unit Topics

1) Water of Crystallisation

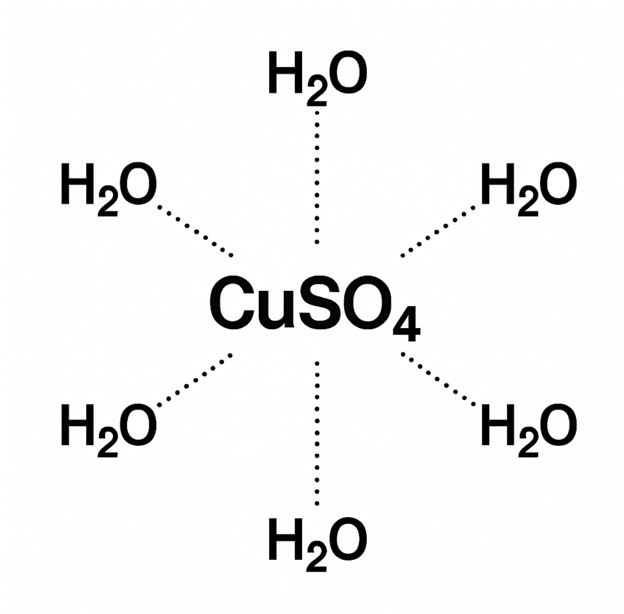
2) Mole Concept + Calculations
3) Diffusion
4) Collision Theory

## Water of Crystallisation

**Certain compounds can absorb water**. These compounds are called **hygroscopic** compounds, as **water molecules surround them**.

The way to write down if a compound is surrounded by water molecules is by adding a dot and then writing a number next to $H_2O$ to denote the number of water molecules.

$CuSO_4$ is a *hygroscopic* compound. It surrounds itself with **6 water molecules**.
$CuSO_4 \cdot 6H_2O$
We added a dot, and then wrote the number of water molecules.
This is how we would draw it:



# Moles

## Calculating Moles

**Mole** is an **SI Unit** which is used to measure the **amount of substance**.
It is the **ratio** of that substance's **mass in grams** and **molar mass**.
Moles = Mass in Grams/Molar Mass
Its **unit is** "**mol**". Just like for mass we have Kg, we have mol for Moles

## Calculating Molar Mass

To find how many moles there are, you need to first calculate the molar mass of the element/compound.
This is done by **adding up the mass numbers of all the elements in the compound**. Atomic mass can be read off of the periodic table, it is the **sum of neutrons and protons**. So, it's basically the **weight** of the compound.
The unit for Molar Mass is "**amu**" (atomic mass unit) *or* "**g/mol**"

Examples:

Calculate the Molar Mass of **NaOH**
First, we separate the compound into the different elements:

1) Na
2) O
3) H

Then we see the periodic table and find all of their mass numbers
Na → 23
O → 16
H → 1
Now we **add them up**: 23+16+1 = **40 amu or 40 g/mol**

Calculate the Molar Mass of $CuSO_4$ **.** $5H_2O$ (if you don't understand why we have a dot here, scroll up and revise "Water of Crystallisation" topic ***<u>again</u>***).
This includes:

1) Cu
2) S
3) 4 x O
4) 5 x $H_2O$

Cu = 64
S = 32
O = 16
H = 1
64 + 32 + (4 x 16) + 5 x (1 x 2 + 16) = 250amu

**Break Down (Explanation)**
We added 64 and 32 for Copper and Sulphur. Then, we multiplied 16 (which is mass number of oxygen) by 4 because **there were 4 oxygen atoms**. Then we did 1 x 2 + 16 in brackets because hydrogen's mass number is 1, and there are 2 hydrogen atoms, so we multiply by 2. Then we add 16, because oxygen is 16. This gives us the molar mass of $H_2O$. Since there are 5 $H_2O$ we multiply by 5 and then add them all up.

Now that we know how to calculate the molar mass, we can calculate the moles.

Mole Question

Calculate the number of moles in 300 grams of Sodium Hydroxide.
Sodium = $Na^+$
Hydroxide = $OH^-$
Sodium Hydroxide = $Na^+ + OH^- \rightarrow NaOH$
Mole = Mass in grams/molar mass

Mass in grams = 300
Molar mass = ???, we must find it

Find all mass numbers first:
Na = 23
O = 16
H = 1
Add them up: 23+16+1 = **40 amu**

Use the formula:
300/40 = 7.5mol

**There are 7.5 moles in 300 grams of sodium hydroxide.**

**NOW TIME FOR THE HARD STUFF**

All Mole Formulae:
1) Moles = Mass(g)/Molar Mass
2) Moles (for gases) = volume in dm/22.7
3) Moles = Number of Particles/Avogadro's Number
4) Number of Particles = Moles x Avogadro's Number

**Avogadro's Number = 6.02 x $10^{23}$**

## Practice Questions:

**Q1) Calculate the number of particles in 500g of Sodium Chloride**
Number of Particles = Moles x Avogadro's Number
We need to find Moles

Moles = mass(g)/molar mass
Molar mass = 23+35.5
Molar mass = 58.5
Moles = 500/58.5
Moles = 8.55
Number of Particles = 8.55 x 6.02 x $10^{23}$

**Number of particles = 5.15 x $10^{24}$**

**Q2) Calculate the number of particles of 72.5 grams of HCl**

Number of Particles = Moles x Avogadro's Number

We need to find Moles

Moles = mass(g)/molar mass

Molar Mass = 1 +35.5

Molar Mass = 36.5

Moles = 72.5/36.5

*Moles = 1.986*

Number of Particles = Moles x Avogadro's Number

Number of Particles = 1.986 x 6.02 x $10^{23}$

**Number of Particles = 1.196 x $10^{24}$**

**Q3) Calculate the mass of $NaHCO_3$ in grams if it has 36.8 x $10^{23}$ particles.**

Moles = Mass(g)/Molar Mass

We need Moles and Molar Mass

Molar Mass = 23+1+12+(16 x 3)

*Molar Mass = 84 amu*

Moles = Number of Particles/Avogadro's Number

Moles = 36.8 x $10^{23}$/6.02 x $10^{23}$

*Moles = 6.113mol*

Moles = Mass(g)/Molar Mass

6.113 = Mass(g)/84

Mass(g) = 6.113 x 84

**Mass(g) = 513.492g**

**Q4) How many particles of $CO_2$ present in 12.2dm^3 of volume.**
Number of Particles = Moles x Avogadro's Number
We need number of Moles

Moles (in a gas) = Volume in dm/22.7
Moles = 12.2/22.7
*Moles = 0.537mol*

Number of Particles = Moles x Avogadro's Number
Number of Particles = 0.537 x 6.02 x $10^{23}$
**Number of Particles = 3.233 x $10^{23}$**

## Chemical Formula Related Mole Questions
These are pretty difficult.

In these questions you are given a chemical formula, in which 1 compound's mass is given from the reactant and you have to find the mass of the product.

There are 4 steps:

1) **Balance the equation**

Just balance it

2) **Find Moles of compound whose mass we are given**

Use formula Moles = Mass(g)/Molar Mass

3) **Develop Molar Ratio**

First make a ratio of the number next to the compound whose mass we are given, and the number next to the reactant.
Then, make another ratio of the moles you found of the compound whose mass we are given, and "x", which will be the moles of the reactant.

4) **Find Unknown**

Using the unknown we can find the mass of the compound

Practice Questions

Q1) B + $O_2$ → $B_2O_3$

(300g) (Excess) (Mass in grams?)

1) Balance Equation

**4**B + 3$O_2$ → **2**$B_2O_3$

2) Find moles of boron

Moles = mass(g)/molar mass

Moles = 300/11

Moles = 27.28mol

3) Develop Molar Ratio

B : $B_2O_3$

In the equation, there are **4 borons** and **2** $B_2O_3$. So the ratio will be 4:2

**4** : **2**

We found the moles of Boron to be 27.28, and we need to find the moles of $B_2O_3$, so we make another ratio, where "x" is the number of moles in $B_2O_3$

27.28: x

So then we have:

B : $B_2O_3$

4 : 2

27.28: x

Now we cross multiply, so 4 times x is equal to 2 times 27.28

4) Find unknown

$(4)(x) = 27.28 \times 2$

x = 13.64 (this is the number of moles in $B_2O_3$)

5) Find mass

Now we plug in the values of the formula. We will have to also find the molar mass, which we can get from the periodic table.

Moles = mass(g)/molar mass

13.64 = mass(g)/ 11 x 2 + 16 x 3 (boron is 11, oxygen is 16)

13.64 x 70 = Mass(g)

**mass(g) = 954.8**

**954.8 grams of** $B_2O_3$ is produced.

Q2) $CaO + CO_2 \rightarrow CaCO_3$

(50g) (Excess) (Mass=?)

Excess is written just because it is idk why i asked sir he said just forget about it.

1) Balance Chemical Equation

The nice thing about this equation is that it is already balanced, so we can go directly to mole calculation:

2) Find moles of CaO

Ca = 40 amu

O = 16 amu

Molar mass = 40+16 = 56

Moles = mass in grams/molar mass

Moles = 50/56

Moles in CaO = 0.893

3) Develop Molar Ratio

$CaO : CaCO_3$
1 : 1
0.893 : x

4) Find Unknown

1 times x is equal to 1 times 0.893
1x = 1 x 0.893
x = 0.893 moles (in $CaCO_3$)

5) Find mass in grams of $CaCO_3$

Moles = mass(g)/molar mass

MOLAR MASS:
Ca = 40 amu
C = 12 amu
$O_3$ = 16 x 3 amu

0.893 = mass(g)/40+12+16x3
0.893 = mass(g)/100
Mass (g) = 0.893 x 100
**Mass = 89.3g**
89.3 grams of Calcium Carbonate is produced.

## Diffusion

Diffusion is the **change of the concentration** of particles from high concentration to low concentration. It is an essential part of chemical change as it **allows particles to move about and collide**, allowing for them to react.

There are **4 main factors** of diffusion:

1) **Temperature**

If the temperature of the particles increases, their kinetic energy also increases. As a result, the particles move faster, allowing them to diffuse more quickly from one place to another.

Example:
There are 2 samples of $NH_3$ (g), one is heated to 80 degrees celsius and one is heated to 100 degrees celsius. Which one would diffuse faster? The one heated to 100 degrees celsius because particles with higher temperature diffuse faster.

2) **Molar Mass**

Molar mass, as discussed previously, is basically the weight of the particles. The heavier the particle, the slower it will be able to move around, and vice versa. Thus, **more molar mass = slower diffusion**.

Example:
Which compound will diffuse faster, $CH_4$ or $NH_3$?
$CH_4$ Molar mass = 12 + 1 x 4 = 16 amu
$NH_3$ Molar mass = 14 + 1 x 3 = 17 amu
Methane is lighter than ammonia, therefore it will diffuse faster.

3) **Concentration**

If the particles are already very concentrated, then they will diffuse faster as there will be a higher chance of them spreading out into less concentrated areas.

4) **Concentration Gradient**

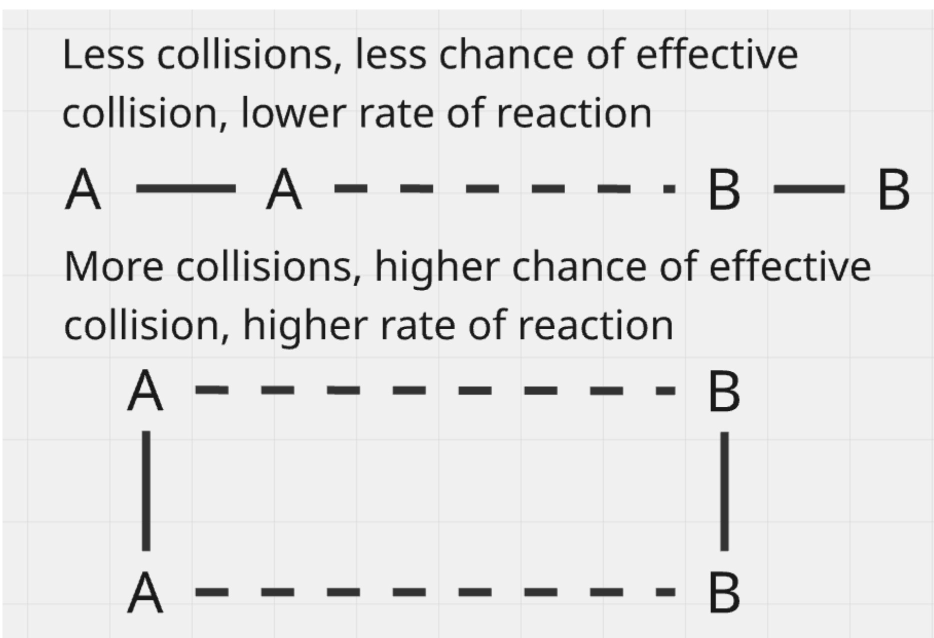
The concentration gradient is the difference in concentration between 2 regions. The larger the concentration gradient, the faster diffusion will occur. For example, if the concentration of 1 region is very high, and the concentration of another region is very low, the concentration gradient will be high. As a result, diffusion from the high concentration to the lower concentration will be faster.

## Collision Theory

Collision theory states that “**In order for substances to react, they must collide”**. Basically, they need to touch each other in order for a reaction to happen. Its factors define the **rate of reaction**. There are **5 factors**.

1) **Arrangement of Particles**

In order for more **effective collisions** to take place i.e collisions that result in a reaction taking place, particles should collide in the correct **orientation and direction**. Furthermore, **more collisions in correct orientation = higher chance of effective collisions**, so if particles are arranged to have more effective collisions then the rate of reaction would also increase.




2) **Temperature**

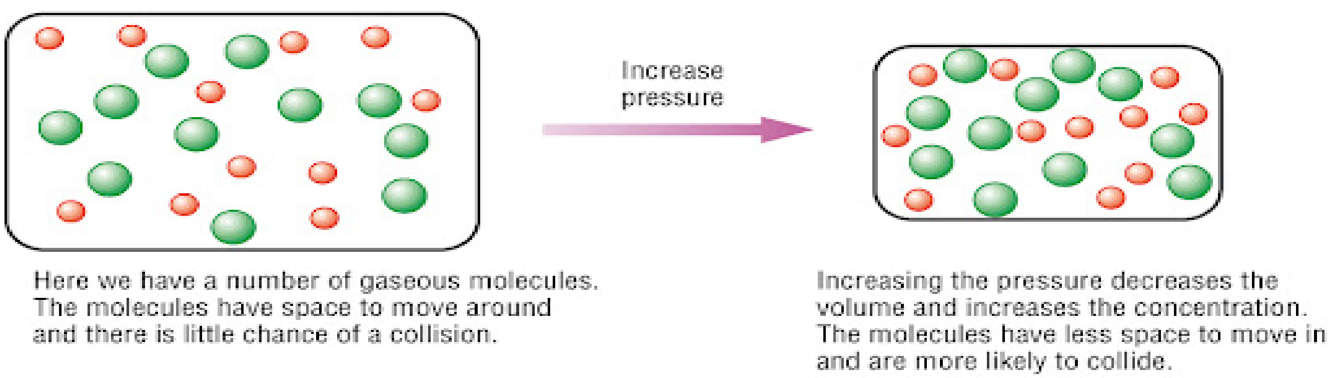
More temperature → more kinetic energy → particles move more → particles collide more → higher chance of an effective collision → higher rate of reaction.

Furthermore, more kinetic energy → more energy in atoms → activation energy better fulfilled (A.E is discussed on the next page).

### 3) Pressure

Higher pressure means that the particles have **less space** to move around without colliding with the wall of the container or with other particles. Thus, if we have higher pressure then particles will **collide more** (as there is less space between them), meaning more collisions → more effective collisions → higher rate of reaction.

**High pressure = Faster rate of reaction**




### 4) Surface Area

An  important thing to remember is that when 2 substances react, only the **particles that are exposed will collide with each other**, all the particles inside the substance will not react. Therefore, if we **increase the surface area**, then **more particles will be exposed**, leading to **more collisions**, higher rate of **effective collisions** and **higher rate of reaction.**

This is also why many small pieces of sugar dissolve in tea quicker than 1 big block of sugar. Even though the total mass is the same, the small particles have **more surface area**, meaning that more particles can collide with the tea at the same time.

### 5) Catalyst

A catalyst is something that **speeds up a reaction** by **reducing the activation energy** required for the reaction to take place. Most **transition metals** are good catalysts.

What is activation energy?
When forming a bond, energy is required for 2 things:

1) To break the bonds the particles initially had with each other
2) To form new bonds between the particles that collided

Activation energy is **the sum of these energies**.
Catalysts reduce both of these by providing an **alternate pathway** which requires **less activation energy**.

In summary:
For atoms to react, they must

1) Collide in the correct orientation, called an **effective collision**.
2) Have enough **energy** (called the activation energy) to break original bonds and create new ones.

All of these factors help with one of these requirements

# Unit 7 – Isotopes

## Unit Topics

1) Isotopes and Their Properties
2) Stability of an Isotope
3) Unstable Isotopes + Half-lifes
   a) Alpha Rays
   b) Beta Rays
   c) Gamma Rays
4) Properties of Rays

# Isotopes and Their Properties

**Elements don't all come in 1 form**, they come in various forms. These forms have different atomic masses, but the same atomic number. Basically, the **number of neutrons changes** with each form.

For example, Hydrogen can have 0 neutrons, 1 neutron or 2 neutrons.

This brings up the question, if these elements have varying neutron numbers, then they will also have different atomic masses (weight), so how do we decide what the atomic weight of an element is?

Through a process called **Mass Spectrometry** we determine the **relative abundance** of each isotope, which is how often they occur in a sample. After that, we **take the average mass** of all of the isotopes to get the **relative atomic mass** which is the average atomic mass of an element.

The chemical properties of an element depend upon the atomic number, whereas the physical properties depend upon the atomic mass.
The atomic number across isotopes of the same element is the same, so the **chemical properties stay the same**. However due to differences in atomic masses (different weights), **physical properties** such as boiling points, melting points, density etc are **different**.

# Stability of an Isotope

Isotopes have different stabilities based upon their properties.

Factors of Stability

1) **Atomic Size**

If an atom has more than 83 (some say 88) protons then it is deemed as unstable as the repulsion forces between them become too high.

2) **Proton to Neutron ratio (protons divided by neutrons)**

If the ratio is close to 1 then it is stable, if it is far off then unstable.
For Example:
Oxygen has 8 protons and 8 neutrons. 8/8 = 1, which means it is very stable

On the other hand, Radium has 138 neutrons and 88 protons. 88/138= 0.638, which is far off from 1, meaning that it is quite unstable.

Iron is the most stable atom.

## Unstable Isotopes

If an isotope is unstable, like all things in nature it will move towards stability. It does this through **decay** in which the atom releases radiation and becomes a **new element**. The time taken for **half of the atom to decay** is the **Half-life.**

***Half-Life:***
- Time for half of an element to decay.

***Represented by:***
$t_{½}$

- **t** represents time
- **½** is subscripted (½ is half)

***Example of Half-Life:***
$_{92}U^{238}$ (100%) ⟶ (50%) ⟶ (25%) … etc

- In this example, Uranium is going through the Half-Life process.

- It keeps dividing by half until it completely decays.

- 50% = 1st Half-Life
- 25% = 2nd Half-Life
- And so on.

Whilst undergoing its half-life, the **element releases radiation rays.** These are:

1) Alpha Rays
2) Beta Rays
3) Gamma Rays

## Properties of Alpha Rays ($\alpha$)

1) **Nature**

They have 2 protons and 2 neutrons, which is the same as helium. So, we say that **alpha rays have a helium nucleus**.
When writing equations we denote alpha rays as $\alpha$, or $He_2{}^4$.

2) **Example Equation**

$${}_{92}U^{238} \longrightarrow {}_{90}Th^{234} + {}^{4}{}_{2}He^{2+}$$

Uranium loses 2 neutrons and 2 protons, so it becomes Thorium. Alpha particles are also released.

3) **Penetration**

Alpha particles are **big**, so they can't penetrate deep.

4) **Ionization** (how reactive it is)

Alpha particles have very **high ionization (**they **react a lot)**.

5) **Transformation**

When released, the **element changes**.

## Properties of Beta Rays (β)

1) **Nature**:

There are **2 types of beta particles**:

1) Electron $\beta^-$ → **Negative Charge**
2) Positron $\beta^+$ → **Positive Charge**

2) **Equations**

**2.1) Beta-minus Equations:**

$$_{92}U^{238} \longrightarrow {}_{93}Np^{238} + \beta^-$$

$$_{90}Th^{238} \longrightarrow {}_{91}Pa^{238} + \beta^-$$

When there are **too many neutrons** in an atom, a neutron is converted into a proton **increasing the atomic mass by 1**. To make up for this a beta-minus particle is released.

**2.2) Beta-plus Equations:**

$$_{92}U^{238} \longrightarrow {}_{91}Pa^{238} + \beta^+$$

When there are **too few neutrons** in an atom, a proton is converted into a neutron **decreasing the atomic mass by 1**. To make up for this a beta-plus particle is released.

3) **Penetration**

Beta particles are **medium weight**, so they have **moderate penetration**.

4) **Ionization** (how reactive it is)

Beta particles have **moderate ionization (react moderately)**.

5) **Transformation**

When released, the **element changes**.

## Properties of Gamma Rays ($\gamma$)

**Nature**

They have **no mass**. Just **energy.** Released **alongside alpha and beta particles** when the atom is in an excited state. This is represented by an *.

**Example Equation**

$${}_{90}Th^{234*} \longrightarrow {}_{90}Th^{234} + \gamma$$

**Penetration**

Gamma particles are **extremely small**, so they penetrate deep.

**Ionization** (how reactive it is)

Gamma particles have very **low ionization (barely react)**.

**Transformation**

When released, the **element doesn't change**.

# Unit 9 – Impacts of the Chemical Industry

Unit Topics

1) Greenhouse Effect

a) Greenhouse Gases

b) Carbon Cycle

2) Global Warming and Climate Change

3) Acid Rain

a) Causes

b) Terrestrial Impacts

c) Aquatic Impacts

4) Ozone Holes/Depletion

5) Algal Blooms

6) Heavy Metals Poisoning

## Greenhouse Effect

### Greenhouse Gases

The greenhouse effect is caused by 3 main greenhouse gases:

1) $CO_2$
2) $SO_4^{-2}$
3) $CH_4$

When **heat radiation** from the sun enters our atmosphere, it **warms** the planet through its presence, bounces off the surface and then exits the atmosphere. **Greenhouse gases create a barrier** in the atmosphere, preventing some of the heat radiation from escaping. This is normally good, since it allows the Earth to **retain heat**. However, too much of these gases result in the Earth **overheating**. This is called the **greenhouse effect**.

## Carbon Cycle

1. What is the Carbon Cycle?

- The carbon cycle describes how **carbon atoms move** through living organisms, the atmosphere, oceans, and the Earth's crust.

- It is essential for life because carbon forms the basis of organic molecules (carbohydrates, proteins, lipids, DNA).

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2. Major Processes in the Carbon Cycle

### a) Photosynthesis

- Plants, algae, and some bacteria take in $\mathbf{CO_2}$ from the atmosphere.

- Using sunlight, they convert $CO_2$ and water into **glucose** and oxygen.

- This process **removes carbon** from the atmosphere.

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b) Respiration

- Living organisms (plants, animals, decomposers) break down glucose for energy.

- This releases **$CO_2$ back into the atmosphere**.

- Equation

$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + \text{Energy}$

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c) Decomposition

- Dead plants and animals are broken down by decomposers (bacteria, fungi).

- Carbon in their bodies is converted back into **$CO_2$** (or methane, $CH_4$ in anaerobic conditions).

- Returns carbon to the air and soil.

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d) Combustion

- Burning fossil fuels (coal, oil, gas) or biomass releases stored carbon as **$CO_2$**.

- Rapidly increases atmospheric $CO_2$ compared to natural processes.

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e) Fossilization

- Over millions of years, buried dead organisms may turn into **fossil fuels** (coal, oil, gas).

- Stores carbon in the Earth's crust

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3. Human Impact

- **Burning fossil fuels:** adds extra $CO_2$ → enhances greenhouse effect → global warming.

- **Deforestation:** reduces photosynthesis → less $CO_2$ removed from atmosphere.

- **Ocean acidification:** more $CO_2$ absorbed → harms marine ecosystems.

## Global Warming and Climate Change

Global warming has many affects:

1) Increase in temperature results in **ice caps melting** as water has more volume than ice. This results in **loss of habitat** and **flooding**.
2) Due to increased evaporation, regions become more **dry**, making the frequency of **droughts** increase. This makes them devoid of nitrogen (essential for proteins) and fertility.
   Furthermore, to regulate the temperature the Earth starts to rain, resulting in **flooding**.
3) Extreme dryness also results in **forest fires**, causing habitat loss and deforestation, which **increases carbon emissions**.

All of the mentioned effects come together to form **climate change**.

## Acid Rain

Causes

There is a lot of $CO_2$ in the environment, so water reacts with it to form $H_2CO_3$.

There are many **oxides in the air**, such as $SO_2$ and $NO_2$. To regulate them, the earth uses water to protect itself. However, these **react to form acids**:

$SO_2 + H_2O \rightarrow H_2SO_4$

$NO_2 + H_2O \rightarrow HNO_3$

These acids cause **acid rain.**

Impacts of Acid Rain:

1) Terrestrial (Land)

a) **Infrastructure:** Acid reacts with carbonates (what historical sites are made out of) to form water, carbon dioxide and salt, causing **erosion**.

b) **Living organisms:** When acid rain comes into contact with humans, it causes **respiratory problems, eye infections** and **skin issues**.
Also, the pH of the soil matters a lot for plant growth as enzymes depend on them. If it changes, **fertility is ruined**.

2) Aquatic (Water)

a) **Loss of drinkable water:** When coming into contact with water the pH becomes acidic, making it undrinkable.

b) **Loss of marine life:** Sea animals will die if the water becomes too acidic.

## Ozone Holes/Depletion

When **CFCs** reach the upper atmosphere, **UV rays break them apart**,

releasing **chlorine atoms**.

These chlorine atoms react with ozone ($O_3$), breaking it down into oxygen

($O_2$).

Since chlorine is **not used up** in the reaction, one atom can destroy many ozone molecules.

This ongoing breakdown of ozone is called **ozone depletion**.

THING TO KNOW BEFORE DOING ALGAL BLOOMS AND HEAVY METAL POISONING:
Industries that manufacture goods produce bio products, which make them look bad. To make themselves look good they throw it into the sea where authorities and the public can't find it.

## Algal Blooms

Bio products are **rich in nitrogen**, which is a main component in proteins. When dumped in water, the algae synthesize with the proteins resulting in an **overgrowth of algae**. This is called an **algal bloom**, it stays on the surface level of the water and prevents it from **absorbing sunlight**, resulting in plants and fish to die.

## Heavy Metals Poisoning

Heavy metals are a byproduct of chemicals used for many products. They include Mercury, Cadmium and Lead. When dumped, the **aquatic life absorbs them,** and they **become a part of the crops** and **fish**. When we eat them, we get **heavy metal poisoning.**

YOU MUST MEMORIZE ALL THIS:

**RATA:**

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**Molecular Ions**

1. Hydroxyl/Hydroxide ($OH^-$)
2. Nitrate ($NO_3^-$)
3. Cyanide ($CN^-$)
4. Sulphate ($SO_4^{-2}$)
5. Carbonate ($CO_3^{-2}$)
6. Phosphate ($PO_4^{-3}$)
7. Ammonium ($NH_4^+$)

**Commercial Names**

1. H**Cl** (Hydro**chloric** acid)
2. $H_2\mathbf{SO_4}$ (**Sulphuric** acid)
3. $H_3\mathbf{PO_4}$ (**Phosphoric** acid)
4. $H_2\mathbf{CO_3}$ (**Carbonic** acid)
5. $H\mathbf{NO_3}$ (**Nitric** acid)
6. $H_2O$ (Water)
7. $NH_3$ (Ammonia)
8. $CH_4$ (Methane)
9. NaCl (Edible salt)
10. $Ca(OH)_2$ (Lime Water)

Just for general knowledge you can remember glucose (optional) : $C_6H_{12}O_6$

A Carbonate → Compound with $CO_3^{-2}$
A Sulphate → Compound with $SO_4^{-2}$
A Nitrate → Compound with $NO_3^-$

**Diatomic Elements:**

Always in pairs of 2 (need to rata for chemical equation)

BrINClHOF

1. $Br_2$ (Bromium)
2. $I_2$ (Iodine)
3. $N_2$ (Nitrogen)
4. $Cl_2$ (Chlorine)
5. $H_2$ (Hydrogen)
6. $O_2$ (Oxygen)
7. $F_2$ (Fluorine)

### Physical Properties of Acids and Alkalis

Acids

1. Sour
2. Corrosive
3. Conduct Electricity
4. PH Level less than 7 (Less PH = Strong Acid)
5. Red Litmus

Alkalis

1. Bitter
2. Corrosive
3. Conduct Electricity
4. PH Level more than 7 (More PH = Strong Alkali)
5. Blue Litmus

### Chemical Properties of Acids

Acid + Carbonate → Salt + Water + Carbon Dioxide
Acid + Metal → Salt + Hydrogen
Acid + Alkali → Salt + Water

## General Prediction

### Combustion Reaction

Carbon Compound + Oxygen —heat→ Heat + Water + Carbon Dioxide
Example carbon compound: $CH_4$

**Oxides**
Metal Oxide + Water → Metal Hydroxide

Non-metal Oxide + Water → Non-metal acidic commercial name
Example non metal oxide: $CO_2$ (**carbon** is the nonmetal). This non-metal acidic commercial name will be **Carbonic** Acid then ($H_2CO_3$).

Lime Water + Carbon Dioxide → $CaCO_3 + H_2O$
Lime Water: $Ca(OH)_2$
$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$

**Redox Reactions**

| | **Reduction** | **Oxidation** |
|---|---|---|
| **1** | Lose Oxygen | Gain Oxygen |
| **2** | Gain Hydrogen | Lose Hydrogen |
| **3** | Gain Electrons | Lose Electrons |
| **4** | Reduced Oxidation State | Increased Oxidation State |

**Properties Of Metals**

1. **Hard** - it can resist scratching, not easily bent broken or pierced
2. **Ductile** - Can be made into thin wires
3. **Malleable** - Can be hammered into different shapes without breaking
4. **Lustrous** - They are shiny

5. **Sonorous** - They make a sound when struck
6. **High Melting and Boiling Points** - self-explanatory
7. **Electric Conductivity** - Electricity can pass through it easily
8. **Heat Conductivity** - Heat passes and spreads through it fast due to a free electron which helps speed up the process
9. **High Density** - Very compact, this contributes to its hardness
10. **Magnetic Attraction** (sometimes) - some metals attract other metallic objects towards them

The reason for these properties are explained in unit 4 metallic bonding topic.

**Mole Formulae**

Avogadro's Number: $6.02 \times 10^{23}$

Moles = number of particles/avogadro's number

Number of particles = moles x avogadro's number

Moles = mass in grams/molar mass

Moles (for gas) = volume of gas/22.7

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