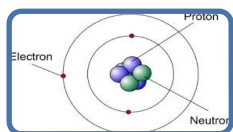
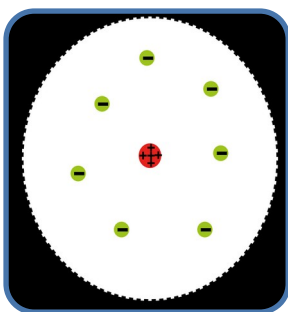
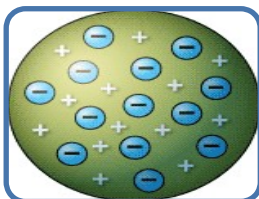


UNIT 1 & 2 NOTES

ATOMIC STRUCTURE

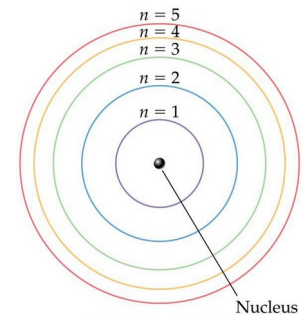
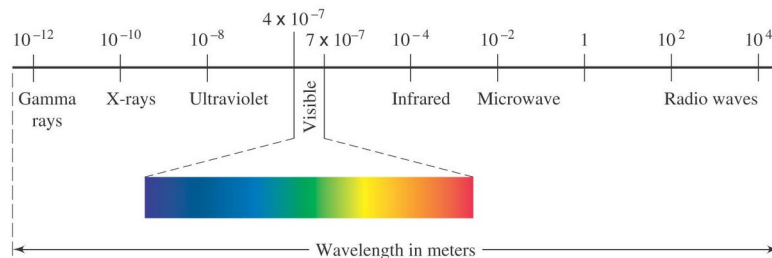
The History of Atomic Structure



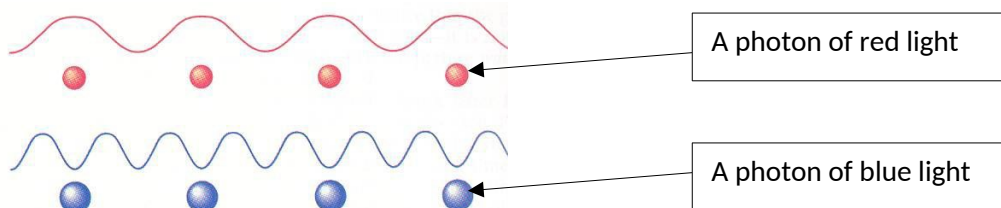
Bohr Model of Atom: Electrons, Flame Tests and AAS

Bohr Model: electrons travel in orbits, or energy levels, around the nucleus. The further the electron is from the nucleus; the more energy it has.

Electromagnetic Spectrum: the classification of electromagnetic waves according to their frequency.



Photons: particles of light energy. Each wavelength of light has photons that have a different amount of energy. The longer the wavelength, the lower the energy of the photons.



Why an element's line spectrum can be used to identify it?

⇒ When atoms are energized, their 'excited' electrons jump to a higher energy level. Because of the instability, the electrons will jump back down to its ground state by releasing the excess energy in the form of light. The wavelengths of the emitted light depend on the difference in energy between the ground state and the excited state. An element has its unique amount of energy in each energy level, so it only emits specific wavelengths and light color.

Flame Test: an analytical technique that relies on an element's unique emission spectrum to identify its presence in a mixture of compound, often used for metal ions.

- 1) A small sample of unknown compound is placed in a hot Bunsen burner flame.
- 2) Sample vaporizes and the heat of flame excites electrons.
- 3) Electrons return to ground state, emitting light with wavelength characteristic of element.

Limitations: only used for small number of metal ions due to limited source of energy in the flame; unclear results, which can be fixed by using a **spectroscope** to see line emission spectrum of a flame.

Metal Ions	Flame Color
Boron	Bright Green
Barium	Pale Green
Calcium	Red/ Orange
Copper	Blue/ Green
Iron	Gold
Lithium	Red/ Crimson
Potassium	Violet/ Lilac
Strontium	Deep Red
Sodium	Yellow

Atomic Absorption Spectroscopy – both quantitative and qualitative

- 1) Use lamp containing the same element (gaseous state) as the one being tested and emits identical wavelengths of light to be absorbed.
- 2) Sample is **vaporized** by the flame and atoms are separated so they can absorb light from the hollow cathode lamp. Only atoms of the element that we are looking for will absorb the emitted wavelengths from the HCL.
- 3) Light passes through a **monochromator**, which is set to select specific wavelength for analysis by detector.
- 4) Detector measures the amount of light remaining without being absorbed = **absorbance value**, which tells us the quantity of the element being tested.

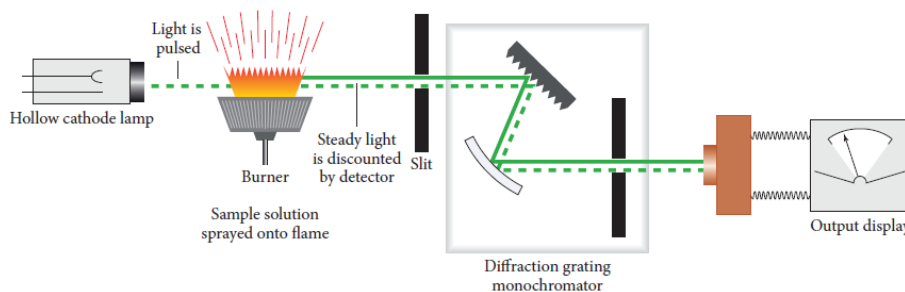


Figure C1.23 The process of atomic absorption spectroscopy

Classifying of Matter

Homogeneous		Heterogeneous	
Pure substance		Mixture	
<ul style="list-style-type: none"> o Particles making up a substance are all of the same kind o Materials with distinct measurable properties (MP, BP, reactivity, strength and density) 		<ul style="list-style-type: none"> o Contain two or more different kinds of substances o Can be composed of elements or compounds or both, which can be separated out by physical properties o Material properties dependent on the identity and relative amounts of the substances that make up the mixture 	
Element	Compound	Solution Ex. Air, sugar dissolved in water solution and glass	Mixture that is not a solution Ex. Nuts mixture or cement
Cannot be decomposed by chemical reactions, either metal or non-metal	Made from two or more elements chemically combined		

Physical Properties	Chemical Properties
Can be found by studying the substance itself rather than its chemical reactions	Describe how it reacts chemically and its tendency to form new substances
Ex. Solubility, state, MP, BP and conductivity	Ex. Reactivity, ability to be stable or decompose with heat and acidity

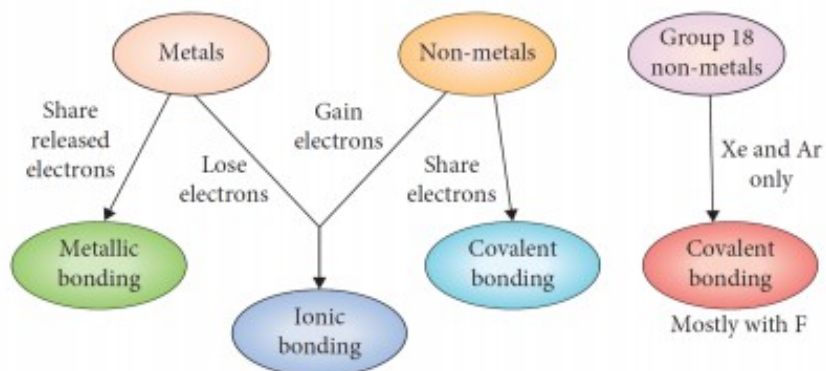
Separation of Mixture

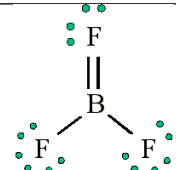
Separation Method	Magnetic	Electrostatic	Sieving	Filtration
Procedure	Remove magnetic materials from a mixture using a magnet	Separate materials that have different charge	Used for mixture of solids or solid a liquid Unlikely to produce pure substance	Separated insoluble solids from a liquid or gas
Physical Properties	magnetism	Electrical charge	Particle size	Particle size and solubility

Decantation	Centrifugation	Separating Funnel	Gravity Separation	Evaporation
Separate solids from a liquid by pouring off the liquid to leave the solid behind	A mixture is spun in a machine called centrifuge . The high speed causes less dense substance to rise to the top	Separates immiscible liquids , which form 2 layers due to difference in density	Heavier particles fall to the bottom when the container is shaken	Heat causes a liquid to turn into a gas, leaving behind any solids that were dissolved in the liquid, as well as impurities
Density	Density	Density and solubility	Force of gravity	Boiling points

Crystallization	(Fractional) Distillation	Chromatography
Purify impure solid substance 1) Dissolve both solid and impurities in solvent at high temperature. Hot filtration to remove any solid impurity 2) Cool to room temp. Desired solid is less soluble at room temp, so it will crystallize 3) Crystals are filtered out of solution, washed with cold solvent and dried.	Separate a solid from a liquid or a mixture of liquids. 1) Solution is boiled, solvent changes to vapor. 2) Vapor passes down a condenser, where it is cooled and converted back to liquid. 3) The liquid is collected as the distillate in the flask.	
Solubility	Boiling point	Intermolecular Forces

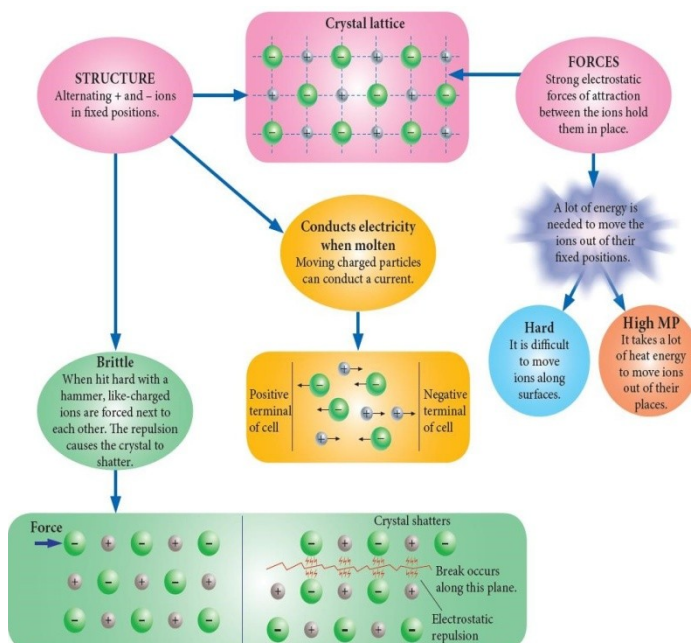
Chemical Bonding



	Giant Lattice			Molecular
	Metallic	Ionic	Covalent Network	Molecular Covalent
Components	Metals	Metals and Non-metals	Group IV Non-metals	Non-metals
Examples	Cu, Fe	NaCl, CaO	SiO ₂ , C ₆₀	CO ₂ , H ₂ O, Cl ₂
Type of individual particles	Positive ions surrounded by negative delocalized electrons	Positive and negative ions	Atoms	Small individual molecules
Bonding Force	Strong electrostatic attraction	Strong electrostatic attraction	Strong covalent bonding due to strong electronegativity	Electronegativity Strong intramolecular covalent bonding Weak intermolecular forces
Directional Bonding	✗	✗	✓	✓
Delocalized Electrons	✓	✗	✗ Except for graphite	✗
Model	Mobile sea of electrons	Ionic lattice	Continuous array (diamond) Graphene (graphite) Soccer ball shape (fullerenes)	Clusters of molecules
Lewis Dot Diagram	✗	$[\text{Mg}]^{+2} [\text{:Br:}]^{-1}$	✗	

Ions with a transition metal = colored

Ions with an alkaline/ alkaline earth metal = white



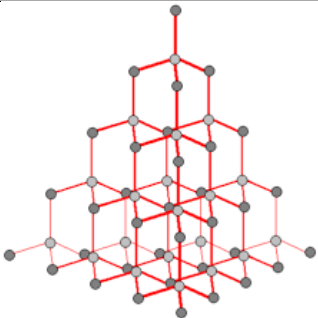
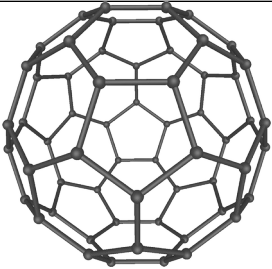
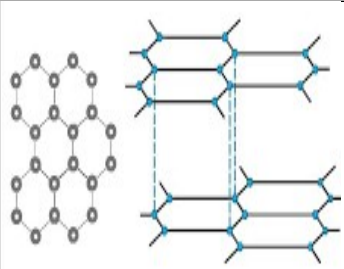
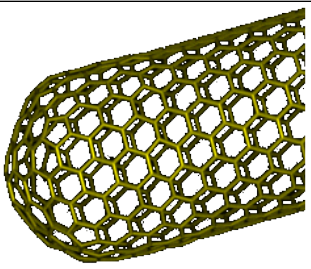
Properties of Types of Bonding

Properties	Metallic	Ionic	Covalent	
			Molecular	Network
Metallic characteristic (reactivity/ tendency to lose electrons)	Very high	Moderate	Low	Low
M.P & B.P	High ⇒ Strong electrostatic attraction ⇒ The more valence electrons, the higher MP and BP	High ⇒ Strong electrostatic attraction between cations and anions	Low ⇒ Weak inter. Forces	Very High ⇒ Strong intra. Molecular covalent bonding lattice
Hardness	Generally hard ⇒ Strong bonding ⇒ Close packing ions	Hard ⇒ Strong bonding	Soft ⇒ Weak inter. Forces	Very hard ⇒ Strong lattice (except for graphite)
Brittleness	Not brittle ⇒ Non- directional bonding so electrons can slide over ions	Very brittle ⇒ Directional bonding, shifted ions cause like charged ions to repel	Soft	Brittle ⇒ Directional bonding, once disrupted the lattice is brittle
Malleability (hammered into sheets) & ductility (drawn into wires)	Malleable & ductile ⇒ Flexible delocalized electrons	None, brittle	None, soft	None, brittle
Solubility in water	Insoluble	Soluble ⇒ Ion-dipole forces pull ions from its lattice	Insoluble ⇒ No attraction between water and the molecules	Insoluble
Electrical conductivity	Good ⇒ Delocalized electrons	Only conductive in liquid and aqueous phases ⇒ Charged ions	Insulator ⇒ Localized electrons	Insulator (except for graphite)
Thermal conductivity	Good ⇒ Delocalized electrons and vibration of ions	Quite good ⇒ Vibrations of ions, but slow	Poor	Poor (except for graphite)

Allotropes of Carbon

Allotropes:

- different forms of the same element;
- have similar chemical properties;
- but different physical properties due to different arrangement/ structure.

Structure			
Diamond	Fullerenes	Graphite	Carbon nanotubes
⇒ Rigid 3-D tetrahedral structure ⇒ Each atom is covalently bonded to 4 other atoms ⇒ No free electrons ⇒ Very strong covalent bonding	⇒ 60 ⁺ carbon atoms ⇒ Arranged in a sphere or cage ⇒ Each atom is bonded to 3 other atoms ⇒ 1 delocalized electron/ atom ⇒ Weak inter. forces ⇒ Strong intra. forces	⇒ Layers of 2-D hexagonal shaped rings. ⇒ 6 rings (graphene)/ layer ⇒ Each atom is bonded to 3 other atoms ⇒ 1 delocalized electron/ atom ⇒ Weak inter. ⇒ Strong intra.	⇒ Walls made of 1 single layer of graphite rolled into a cylinder shape ⇒ Each atom is bonded to 3 other atoms ⇒ 1 delocalized electron/ atom ⇒ $1^{+1} \text{ dimension} = 1 \text{ nm} = 1 \times 10^{-9} \text{ m}$
			

Conductor of heat & electricity			
Poor ⇒ No mobile electrons	Semi ⇒ Delocalized electrons can't move between molecules	Good ⇒ Delocalized electrons can move between graphene sheets	Good
Hard/ soft & brittle			
Hard & brittle ⇒ Continuous, strong array	Hard & brittle	Soft & tensile strength ⇒ Graphene sheets can slip over each other ⇒ Strong intra.	Same as graphite
BP & MP			
High	High	High	High

Chemistry Calculations

Relative Atomic Mass A_r : mass of an atom compared with one-twelfth of the mass of an atom of carbon-12.

$$A_r = \frac{\%a \times Ar_a + \%b \times Ar_b}{100}$$

Relative Molecular Mass M_r : mass of a molecule compared with one-twelfth of the mass of an atom of carbon-12.

$$A_r = \sum \text{of } Ar \text{ of the component atoms}$$

$$\text{Percentage composition} = \frac{\text{mass of a component}}{\text{mass of the total formula}} \times 100$$

Energy Changes in Chemical Systems

- **Enthalpy/ Heat content (H):** sum of potential and kinetic energy of a substance.
- **Change in enthalpy:** smaller in physical changes than chemical reactions

$$\Delta H = H(\text{products}) - H(\text{reactants}) \text{ J mol}^{-1} / \text{kJ mol}^{-1}$$

Exothermic Reactions	Endothermic Reactions
Release energy and heat up surroundings	Absorb energy and cools the surrounding down
Negative ΔH: Enthalpy of product is reduced	Positive ΔH: Enthalpy of product is increased
<p>Energy Profile Diagram</p> <p> $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) \quad \Delta H = -572 \text{ kJ}$ $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2\text{H}_2\text{O}(\text{l}) + 572 \text{ kJ}$ </p>	<p>Energy Profile Diagram</p> <p> $\text{CO}_2(\text{g}) \longrightarrow \text{C}(\text{s}) + \text{O}_2(\text{g}) \quad \Delta H = +394 \text{ kJ}$ $\text{CO}_2(\text{g}) + 394 \text{ kJ} \longrightarrow \text{C}(\text{s}) + \text{O}_2(\text{g})$ </p>

Biofuels

- **Biofuels:** fuels made from organic materials & are considered “renewable”.

Biofuel	Bioethanol	Biodiesel
How it's made	Produced by fermentation of starchy crops (sugar cane, corn, wheat)	Made by processing feedstock: vegetable oil, soybean oil, animal fats.
Advantages	Few modifications are needed Carbon neutral, biodegradable and non-toxic Suited to both small-scale home production or large scale industrial	

Disadvantages	Carbon dioxide is still produced due to transportation of feedstock, biofuel production and distribution using fossil fuels Deforestation to plant feedstock for resources
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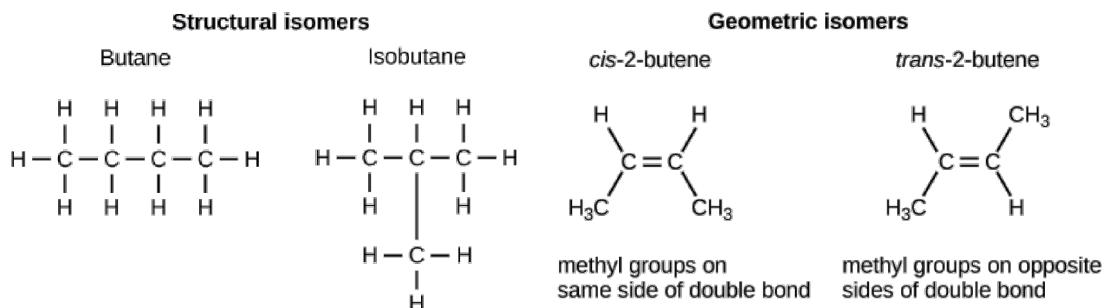
Organic Compounds

- **An organic compound:** made from carbon atoms.
- **Structure & Properties:**
 - covalent bonds
 - low MP & BP
 - flammable
 - soluble in non-polar solvents
 - insoluble in water
- **Hydrocarbons:** organic compounds containing only carbon and hydrogen bonded covalently.
- **4 classes of hydrocarbons:** alkanes-**saturated**, alkenes, alkynes & **aromatic compounds** (benzene) – **unsaturated**.

1. Alkanes: C_nH_{2n+2}

Alkane	Methane	Ethane	Propane
Molecular formula	CH_4	C_2H_6	C_3H_8
Structural formulas			
Expanded	$ \begin{array}{c} H \\ \\ H - C - H \\ \\ H \end{array} $	$ \begin{array}{cc} H & H \\ & \\ H - C & - C - H \\ & \\ H & H \end{array} $	$ \begin{array}{ccc} H & H & H \\ & & \\ H - C & - C & - C - H \\ & & \\ H & H & H \end{array} $
Condensed	CH_4	CH_3-CH_3	$CH_3-CH_2-CH_3$ or $ \begin{array}{c} CH_3 \\ \\ CH_2 \\ \\ CH_3 \end{array} $

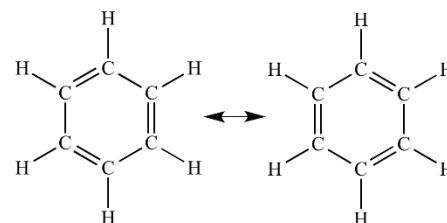
- **Empirical formula:** ratio between C and H
Ex. $C_4H_{10} = C_2H_5$
- **Naming:**
 - prefix: meth-, eth-, prop-, but-, pent-, hex-, hept-, oct-, non-, dec-
 - suffix: -ane
- **State:**
 - 1 – 4 carbon atoms: gas
 - 5 – 8 carbon atoms: liquid
 - 9 – 17 carbon atoms: thick liquid
 - 18+: solid
- **Alkenes:** C_nH_{2n}
- **Isomers:** same molecular formula but different atom arrangements; similar chemical properties and different physical properties.



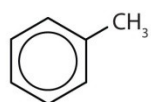
- Alkynes: C_nH_{2n-2}

- Benzene: C_6H_6

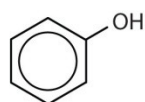
- Structure: Stable, flat hexagonal rings; Identical bonds, which are intermediate in length; C atoms have double bonds to either of its neighboring C atom.



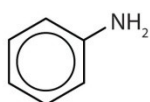
Some benzene molecules can be carcinogen.



Toluene



Phenol

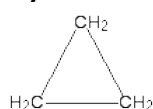


Aniline

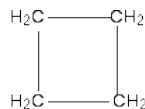
- Aromatic:** Benzene based compounds but one or more H atoms are replaced with other halides/ halogens or alkyl groups.

Ex. 1-bromo- 3- cloro- 5- methyl benzene

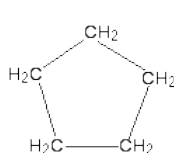
- Cyclic Hydrocarbons:** alkanes, alkenes or alkynes that are arranged in rings; have 2 less hydrogen atoms than non - cyclic forms; **prefix: cyclo-**



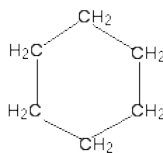
cyclopropane



cyclobutane



cyclopentane



cyclohexane

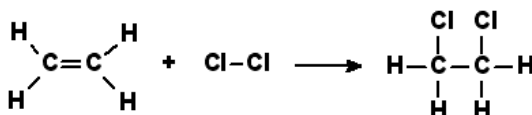
- IUPAC nomenclature for hydrocarbons:**

- Stem name is equivalent to the **longest chain** that contains the **double/ triple bond**.
- Principal functional groups** have to obtain the **lowest number**.
Priorities: double/ triple bond, halogens (F, Cl, Br, I), alkyl groups (methyl, ethyl, ...)
- Prefixes determined by the principal functional group.
- Use 1 word to name the compound:
Name of each group is started with a number indicating its position.
Alphabetical order is applied when listing the groups.
Di, tri, tetra,... do not affect the order.
- Hyphen separates numbers and words.
Comma separates numbers.
- Number attached group from #2.

Reactions of Hydrocarbons

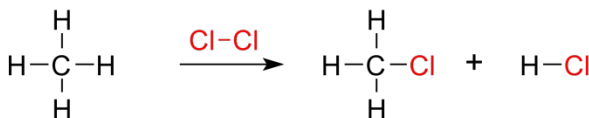
- **Addition Reactions:**

- o Occur with alkene or alkynes (unsaturated reactants).
- o Faster than substitution.
- o Reagents: $\text{H}_{2(\text{g})}$, $\text{Cl}_{2(\text{aq})}$ – chlorine water, $\text{Br}_{2(\text{aq})}$ – bromine water, HCl , HBr , HI .
- o **Observations: Yellow Cl_2 water quickly turns colourless/ Red-brown Br_2 water quickly turns colourless.**



- **Substitution Reactions:**

- o Occur with saturated hydrocarbons (alkanes, benzene and aromatics).
- o Reagents: chlorine water, bromine water, **catalyzed by UV, Pt, Ni and/or heat.**
- o Slower than addition, which is why it is used to distinguish between saturated and unsaturated compounds.



- **Combustion Reactions:**

- o Products: $\text{CO}_{2(\text{g})}$ & $\text{H}_2\text{O}_{(\text{g})}$ & energy (complete combustion)/ $\text{CO}_{(\text{g})}$ / $\text{C}_{(\text{s})}$ soot + $\text{H}_2\text{O}_{(\text{g})}$ & less energy (incomplete combustion).

Ex. Complete combustion: $2\text{C}_4\text{H}_{10(\text{g})} + 13\text{O}_{2(\text{g})} \longrightarrow 8\text{CO}_{2(\text{g})} + 10\text{H}_2\text{O}_{(\text{g})} + 5754 \text{ kJ}$

Incomplete combustion: $2\text{C}_4\text{H}_{10(\text{g})} + 9\text{O}_{2(\text{g})} \longrightarrow 8\text{CO}_{(\text{g})} + 10\text{H}_2\text{O}_{(\text{g})} + 3490 \text{ kJ}$

Rates of Reaction

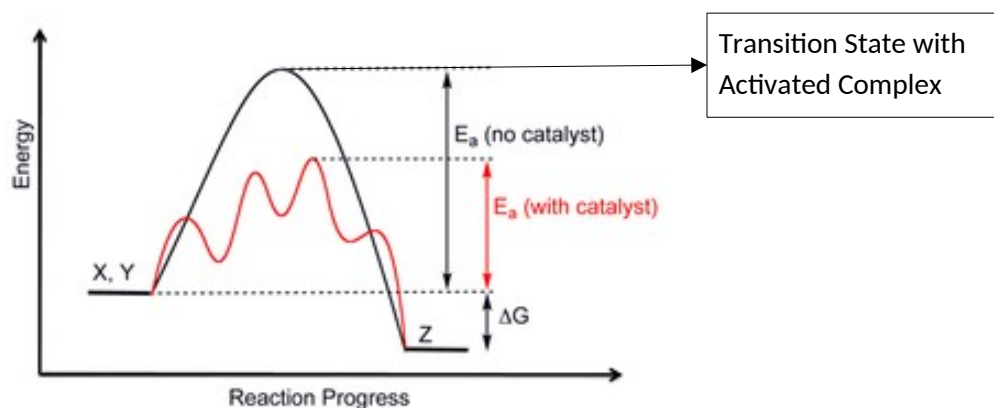
- **Rate of reaction:** rate at which reactants are used up or the rate at which products are formed; measured in mol s^{-1} , g s^{-1} or mL s^{-1} .
- **Collision Theory:** conditions for reactions to occur:
 1. Individual particles of the reacting substances **must collide**.
 2. The collision energy must be **equal to or greater than activation energy**, E_a .
 3. The reacting particles must collide with a **suitable orientation**.

⇒ Increased chance of collisions, increased rate of reaction.

- **Factors of ROR:**

- 1) Nature of reactants: ionic reactions are faster than molecular ones because they have less bonds to be broken and formed as the ions are held together by electrostatic forces of attraction.
- 2) Conc. of reactants: proportional relationship
- 3) State of subdivision of reactants: more surface area, faster ROR.
- 4) Temperature:

- Higher kinetic energy, faster velocity and therefore more collisions.
 - More particles have sufficient activation energy, so more reactions can occur.
- 5) Gas pressure: decreased volume or adding more gas particles increase the pressure, only applied to gas as liquid and solid usually can not be compressed.
- 6) Catalyst: not consumed in a reaction but provide an alternative easier pathway by lowering the E_a .
- Activation Energy**: energy that colliding particles must have to form an activated complex.
 - Transition state**: very unstable point of the reaction where bonds are breaking and forming.
 - Energy Profile Diagram**:



- Examples of catalysts:
 - Enzymes: not consumed, highly specific, lock & key model, specific temperature and pH.
 - Transition metals and their compounds: Pt, Pd, Au, MnO_2 & Rh.
 - Nanoparticles: at least one dimension is 1 – 100nm, have a large surface area.
 - Catalytic converters: reduce toxic gases using platinum, Rhodium, palladium with high SA, increased ROR of CO & unburnt fuel with oxygen, so only carbon dioxide and water are produced; convert NO to N_2 and O_2 .

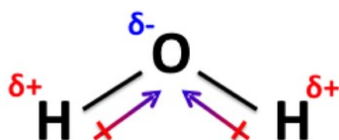
VSPER Theory & Polarity

- Valence Shell Electron Pair Repulsion theory states that:**
 - Pairs of outer shell electrons in atoms form charged clouds which are roughly spherical in shape.
 - These charged clouds repel each other and so are positioned as far apart as possible. This includes both bonding and non-bonding pairs of electrons.
 - However, **lone pairs have larger repulsive forces**.

Shape	Linear	Triangular Planar	Tetrahedral (trigonal planar)	Pyramidal	Bent (V-shaped)
Bond Angle	180°	120°	109°	107°	104.5°
No. bond pairs	2	3	4	4	4
No. lone pairs	-	-	-	1	2

Shape diagram					
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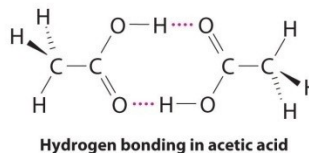
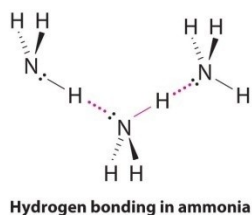
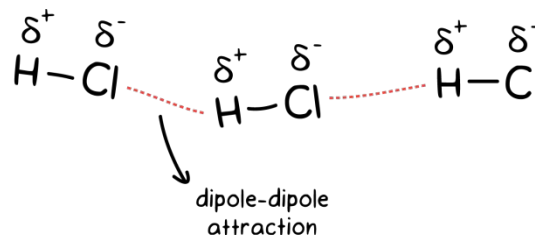
- **Electronegativity:**
 - Attraction of an atom for shared electrons
 - Increases from left to right
 - Highest: fluorine, lowest: lithium.
- **Non-polar covalent bonds:** only between non-metals, EN dif.= 0 to 0.4
 - **Polar covalent bonds:** only between non-metals, EN dif.= 0.5 to 1.7
 - **Ionic bonds:** between metals and non-metals, electrons are transferred, and EN dif.= 1.8+



- **Polar dipoles:** exist in polar bonds/ molecules, negative end attracts more electrons.

Intermolecular Forces

- **Dispersion Forces:**
 - o Weakest intermolecular forces
 - o Exist between all molecules
 - o Temporary instantaneous dipoles
 - o Increases with increased molecular mass and in linear shaped molecules.
- **Dipole-dipole Forces:**
 - o Weak but permanent
 - o Between polar molecules only
- **Hydrogen Bonding Forces:**
 - o Strongest
 - o Between polar molecules
 - o Between the hydrogen atom of a O-H, N-H or F-H bond to the N, O or F atom of another molecule, which has a lone pair that attracts hydrogen atoms.



Generally, network covalent > metallic > ionic > hydrogen bonding forces > dipole-dipole forces > dispersion

Use intermolecular forces to compare & explain physical properties of covalent molecules

- **M.P & B.P:** increase with increased strength of intermolecular forces

- **Vapour Pressure:** tendency of a substance to evaporate, measured in kPa, Pa or atm; increases with decreased intermolecular forces; when vapour pressure = atmospheric pressure => liquid boils.
- **Solubility:** when a new solution is formed, the new solute – solvent interactions are equal or stronger in strength, so covalent network is insoluble due to the strong intramolecular forces; “like dissolves like”.

Chromatography: paper chromatography, TLC, GC & HPLC

- **Mechanism:**
 - **Stationary phase:** solid or liquid supported in a solid (paper)
 - **Mobile phase:** liquid or a gas (water)
 - If components move quickly = high retention factor => they are strongly attracted to the mobile phase
 - If components move slowly = low retention factor => they have strong interactions onto the stationary phase.

$$\text{Retention Factor } R_f = \frac{\text{distance of component}}{\text{distance of solvent}} \text{ both } \uparrow \text{ the pencil line}$$

- **Paper chromatography:**
 - Stationary phase : polar paper (dipole dipole forces and water)
 - Mobile phase: polar water (all intermolecular forces)
- **Thin Layer Chromatography:**
 - Stationary phase: thin layer of silica gel SiO_2 , cellulose, starch or alumina Al_2O_3 coated on a sheet of metal, plastic or glass.
 - Mobile phase: non – polar liquid solvent or water
 - Advantages: simple, rapid, inexpensive; small amount of test substance; non – destructive, only involves physical separation.
- **Gas Chromatography:**
 - Sample is vaporised into gas molecules
 - Stationary phase (coiled to increase SA): solid – gas chromatography (solid absorbant)
liquid – gas chromatography (liquid on an inert support)
 - Mobile phase (carrier gas): inert non-polar gas (He or N_2)
 - Polar and/or large molecules stay longer
 - Non – polar volatile components elute (pass through) faster.
- **High Performance Liquid Chromatography**
 - Mobile phase is pushed through stationary phase under high pressure.
 - Stationary phase: solid particles of silica or polymers tightly packed.
 - Mobile phase: liquid solvent.
 - **Normal phase HPLC: polar S.P, non-polar M.P; vice versa.**
 - Advantages:
 - Can analyse compounds that decompose in gas chrom.
 - Faster due to high pressure
 - Small particles of adsorbent material on stationary phase create a large SA.

Properties of Gases and Kinetic Theory of Gases

Kinetic Theory Assumptions:

- 1) Gases consist of tiny particles moving in rapid, random, straight-line motion until they collide with one another or with the container (**Brownian motion**).
- 2) Collisions between particles or with the walls are perfectly **elastic**.
- 3) The **size of the particles are negligible** compared to the size of their container
 - ⇒ The particles have **mass but no volume**
 - ⇒ Distance between particles are larger than their size
- 4) Any **attractive/repulsive forces** between particles are **negligible**
- 5) **Average kinetic energy increases** as temperature increases

Therefore, gases can be **compressed and diffuse**.

Real Gases	Ideal Gases
Have volume	No volume
Particles attract and repel one another	No forces exist between the particles
At temperature and pressure close to phase changes, gas behaviour is affected by intermolecular forces Stronger intermolecular forces, less molar volume	Molar volume = 22.71 L/mol

Properties of gases

- **Pressure (P):** force exerted by gas against the walls of the container, **increases as temp** and velocity increase.
- **Units:** atm, mm Hg, torr, pascal
1 atm = 760 mm Hg = 760 torr = 100 kPa

$$Pressure = \frac{Force}{Area}$$

- **Volume (V):** space occupied by gas, measured in L or mL.
- **Amount (n):** quantities of gas particles, in grams or moles.
- **Temperature:** average kinetic energy of all molecules

$$E_k = \frac{1}{2} m v^2$$

Gas Laws

- **Boyle's Law** - $P \propto \frac{1}{V}$

$$P_1 V_1 = P_2 V_2$$

- **Charles Laws** - $V \propto T$

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

- **Gay - Lussac's Law** - $P \propto T$

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

- **Avogadro's Hypothesis:** at the same temp and pressure, equal volumes of gases contain the same amount of particles

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$

- **Molar Volume:** volume of 1 mole, or 6.22×10^{23} particles of gas.
- **Standard molar volume:** volume of 1 mole of ideal gas at STP, or at 0°C and 100kPa = 22.71 L/mol.

$$n = \frac{V}{22.71} \text{ at STP}$$

- **Ideal Gas Equation:** when conditions are not at STP

$$PV = nRT$$

Pressure in kPa

Volume in L

n in moles

Temperature in Kelvin

- **Combined Gas Law**

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$