Mole concept & Stoichiometry

- Relative atomic mass (of an element) : weighted average mass of one atom of the element compared to $\frac{1}{12}$ the mass of a ^{12}C atom
- Relative isotopic mass (of an isotope) : mass of an atom of the isotope compared to $\frac{1}{12}$ the mass of a ^{12}C atom
- Relative molecular mass (of an element/cpd) : weighted average mass of a molecule of the element / compound compared to $\frac{1}{12}$ the mass of a ^{12}C atom
- Relative formula mass (of a compound) : weighted average mass of a formula unit of the compound compared to $\frac{1}{12}$ the mass of a ^{12}C atom.

- One mole contains exactly 6.02×10^{23} elementary entities
- Empirical formula : simplest formula that shows the relative number of atoms of each element present in a compound
- Molecular formula : formula that shows the actual number of atoms of each element present in one molecule of a compound

Gases

Assumptions of ideal gas :

- volume of ideal gas particles is negligible as compared to the volume of the container
- There are no intermolecular forces of attraction between the particles of an ideal gas

} MOST ** IMPORTANT

- collisions between the particles of an ideal gas are perfectly elastic
- Gas particles are in rapid, continuous and random motion
- Average kinetic energy of the particles is proportional to their absolute temperature

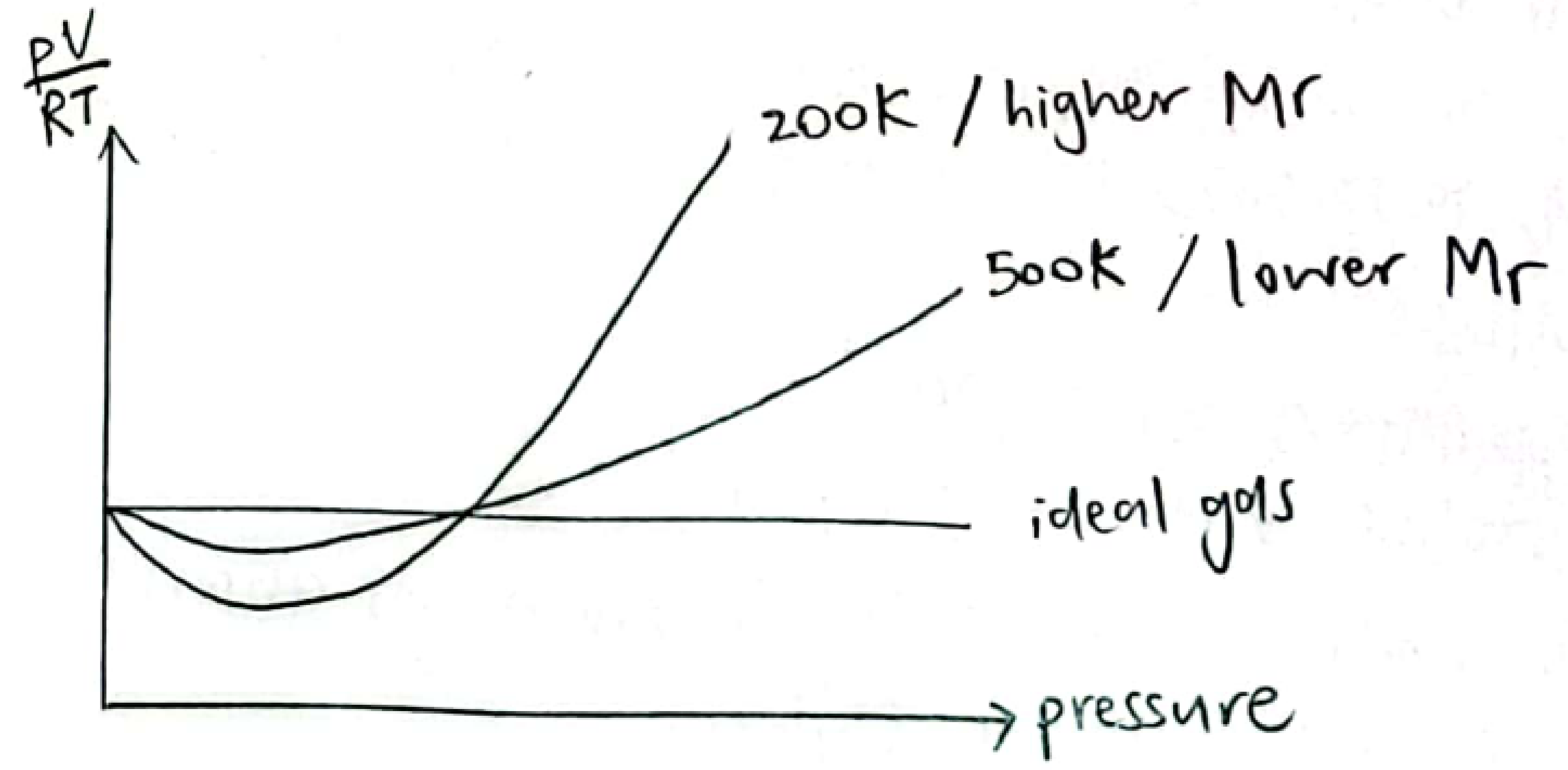
Conditions for more ideal gas :

① Higher temperatures

↳ particles have higher avg. KE → sufficient energy to overcome intermolecular FoA → intermolecular FoA less significant

② Lower pressure

↳ volume of container ↑ → particles are far apart → intermol. FoA less significant → vol. of particles become less significant compared to volume of container



Atomic Structure

- Isotope : atoms of same element with same atomic number but different nucleon number
- Nucleon number : total number of protons and neutrons in the nucleus of an atom

Ionisation Energy across the period :

- no. of protons ↑, nuclear charge ↑
- successive e⁻ added to same outermost shell, shielding effect remains approx. the same
- effective nuclear charge ↑
- net attraction between nucleus and outermost electron ↑
- more energy req. to remove outermost e⁻ → ∴ 1st IE generally ↑ across the period

Exceptions :

- Al smaller 1st IE than Mg (ns^2 vs ns^2np^1)

 Mg: $1s^2 2s^2 2p^6 3s^2$ Al: $1s^2 2s^2 2p^6 3s^2 3p^1$

* · 1st IE of Mg is lower than 1st IE of Al, ∵ less energy req. to remove 3p e⁻ as it is removed from a higher energy 3p subshell

 [∵ 3s orbital lower energy, closer to nucleus → provide some shielding for 3p e⁻] (additional)

- S smaller 1st IE than P

 S: $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$ P: $1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$

 · 1st IE of P lower than 1st IE of S, ∵ less energy req. to remove paired 3px e⁻ in P as it experiences interelectron repulsion

Ionisation Energy Down the Group / Between Periods

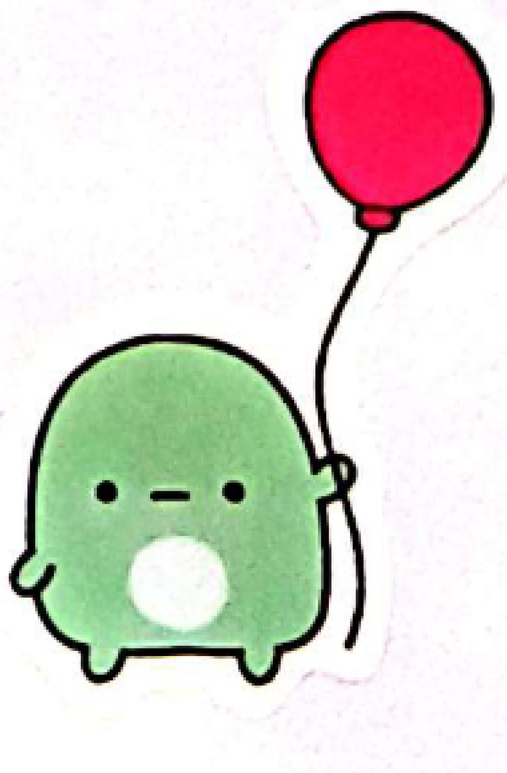
- no. of protons ↑ → nuclear charge ↑
- no. of electron shell ↑ / __ has 1 more electron shell than __ → shielding effect increases significantly
- outermost e⁻ further away from nucleus   *NOTE : no ENC
- net attraction between nucleus and outermost e⁻ ↓
- less energy req. to remove outermost e⁻ → ∴ 1st IE generally ↓ DTG / ↓ significantly from __ to __

Why Successive Ionisation Energy ↑

- no. of protons remain the same → constant positive nuclear charge
- more e⁻ are removed → no. of e⁻ in outermost shell ↓ → shielding effect ↓
- effective nuclear charge ↑
- net attraction between nucleus and outermost e⁻ ↑
- more energy req to remove outermost e⁻ → ∴ successive IE ↑

Atomic / Ionic radius

- same as IE explanation across period
- Atomic / Ionic radius ↓ across the period
- Sharp ↑ in ionic radius from cation to anion, as anion contains 1 more e⁻ shell than cation



Chemical Bonding

- Ionic bond : electrostatic attraction between oppositely charged ions (state ions). Ions are formed by a complete transfer of electrons from a metal to a non-metal
- covalent bond : electrostatic attraction between a shared pair of e^- and positively charged nuclei
- metallic bond : electrostatic attraction between a lattice of positive ions and sea of delocalised valence electrons

VSEPR theory :

- Electron pairs in valence shell of the central atom are located as far apart from one another as possible to minimise repulsion
- Repulsion between electron pairs increase in the order :
  ↳ bond pair - bond pair < bond pair - lone pair < lone pair - lone pair

Strength of bonds

- covalent bond

order of importance

1. Bond order : triple bond > double bond > single bond
   ↳ ∵ of additional side on overlap of p orbitals to form π bond in multiple bonds → stronger bond
2. Atomic radius : shorter atomic radius → shorter bond length → bond stronger
   ↳ ∵ shorter bond length → greater degree of orbital overlap → stronger bond
3. Bond polarity : polar bond stronger than non-polar
   ↳ ∵ polar bonds have additional electrostatic forces of attraction from partial charge
4. Type of hybrid orbital : graphite C–C (sp^2-sp^2) > diamond C–C (sp^3-sp^3)
   ↳ C–C in graphite involve head on overlap of sp^2-sp^2 orbital
   ↳ more s character → greater degree of orbital overlap → stronger bond

- ionic bond
  - Depends on lattice energy, |LE| → more exothermic LE, stronger ionic bond

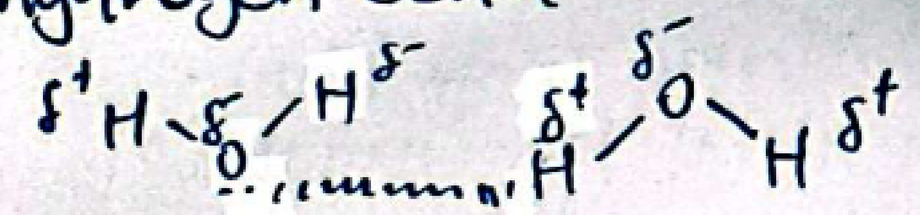
  Factors : $|LE| \propto \frac{|z^+z^-|}{r_+ + r_-}$ } - product of ionic charges
  - interionic radius

- metallic bond
  1. No. of delocalised valence e^- per cation → stronger electrostatic FoA → stronger bond
  2. Size of cation → higher charge density → stronger electrostatic FoA → stronger bond

- id-id
  1. size of electron cloud : greater size → e^- cloud more easily polarised → stronger id-id
  2. shape of molecule :
     ↳ more spherical / branching : less surface area of contact → weaker id-id
     ↳ more elongated / less branching : more surface area of contact → stronger id-id
  3. Packing (for solid state / mp only) → packed more closely → stronger id-id

- pd-pd
  - Polarity of covalent bonds : ↑ polarity of covalent bonds → stronger electrostatic FoA between higher δ^+/δ^- dipoles → stronger pd-pd

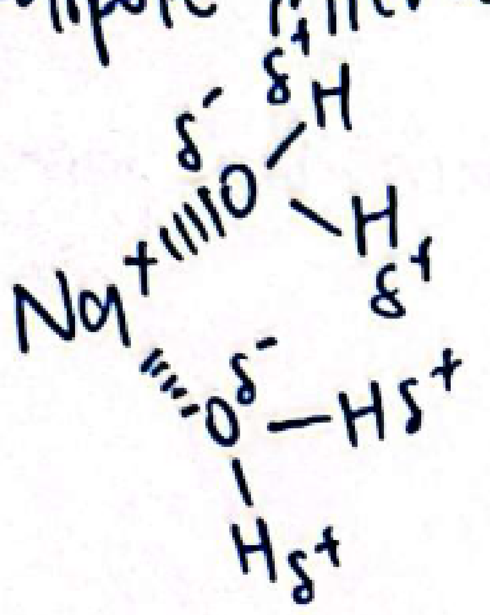
- hydrogen bond

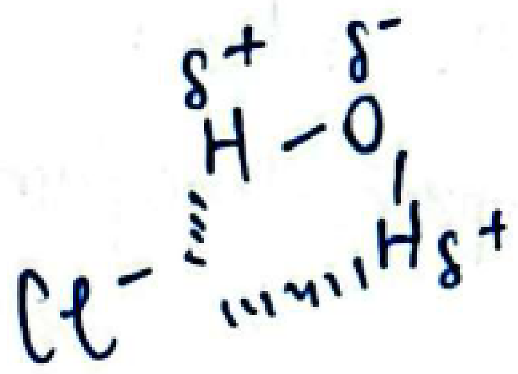


1. Extensiveness of H bonding (avg no. of H bonds per molecule)
   - ↳ larger no. of H bond per molecule
   - ↳ intramolecular H bonding → ↓ extensiveness

2. Electronegativity (F>O>N)

- ion-dipole interaction





- ↑ charge density of ion → stronger
  - ↳ ↓ ionic radius } stronger bond
  - ↳ ↑ ionic charge } stronger bond

Ionic bond w/ covalent character

1. high ionic charge/small ionic radius of CATION
2. ↑ charge density and ∴ polarising power
3. distort/polarise e⁻ cloud of anion to greater exten:
4. introducing a degree of covalent charactor

OR

1. large e⁻ cloud of ANION
2. more easily polarised by cation
3. introducing a degree of covalent character

Physical properties ↗
1. structure
2. bonding
3. particles
4. high/low amt of energy req to overcome ...

- mp/bp

- giant ionic lattice (amino acids/NaCl) → strong ionic bonds between cations & anions
  $|LE| \propto \frac{|z^+ z^-|}{r_+ + r_-}$

- giant metallic lattice (Mg/Na) → strong metallic bond between +vely charged metal cation and sea of delocalised valence electrons

- giant molecular structure (diamond/graphite/SiO_2) → strong covalent bonds between atoms throughout the structure

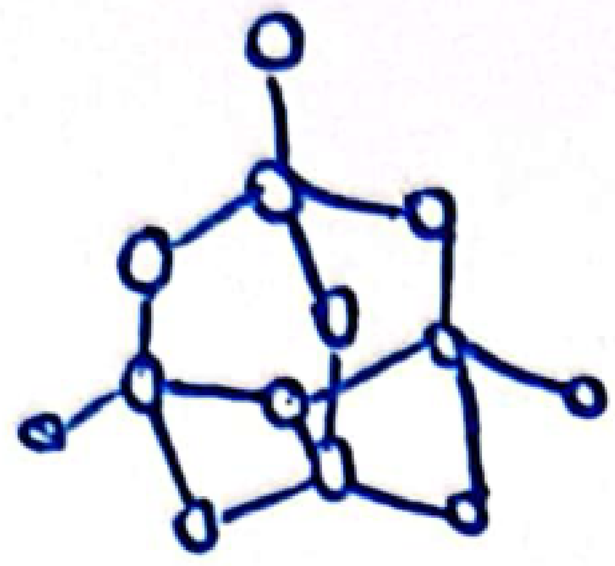
- simple molecular structure (CH_4, NH_3, $AlCl_3$) → -weak id-id / pd-pd forces of attraction between molecules
  -strong hydrogen bonds between molecules

- electrical conductivity
  - presence of mobile charge carriers
    - ↳ there are delocalised valence e⁻ / or mobile ions to act as charge carriers

- solubility

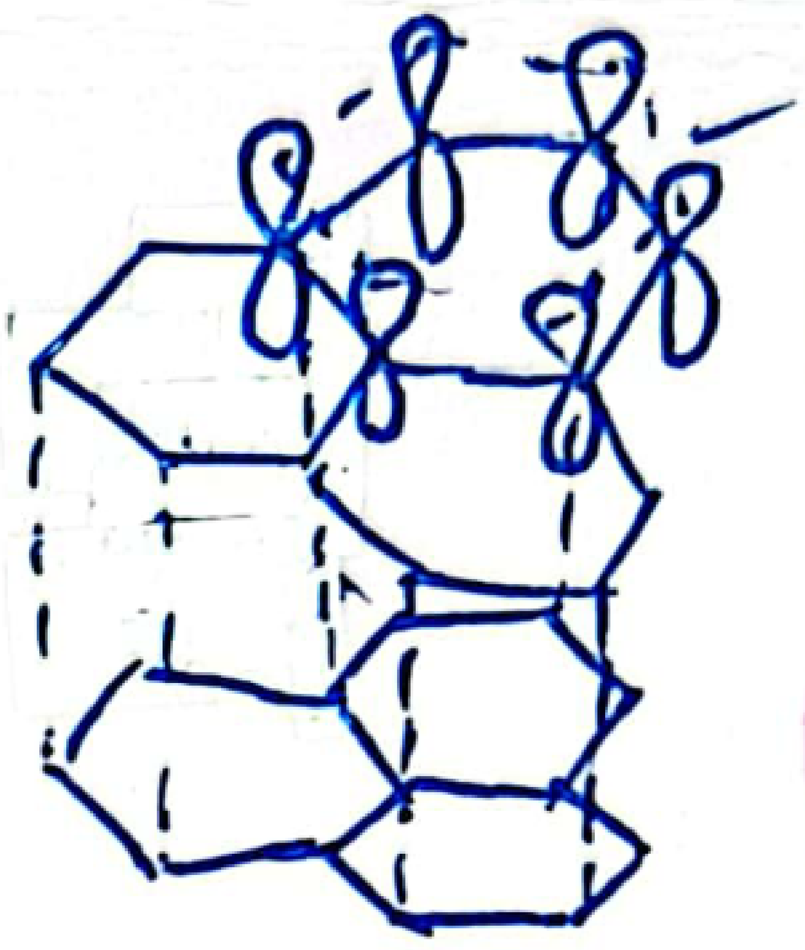
-energy released when <u>(solute)</u> forms _ bond with <u>(solvent)</u>
-provides sufficient / insufficient energy
-to break _ bonds between <u>(solute)</u> and _ bonds between <u>(solvent)</u>

lattice structures



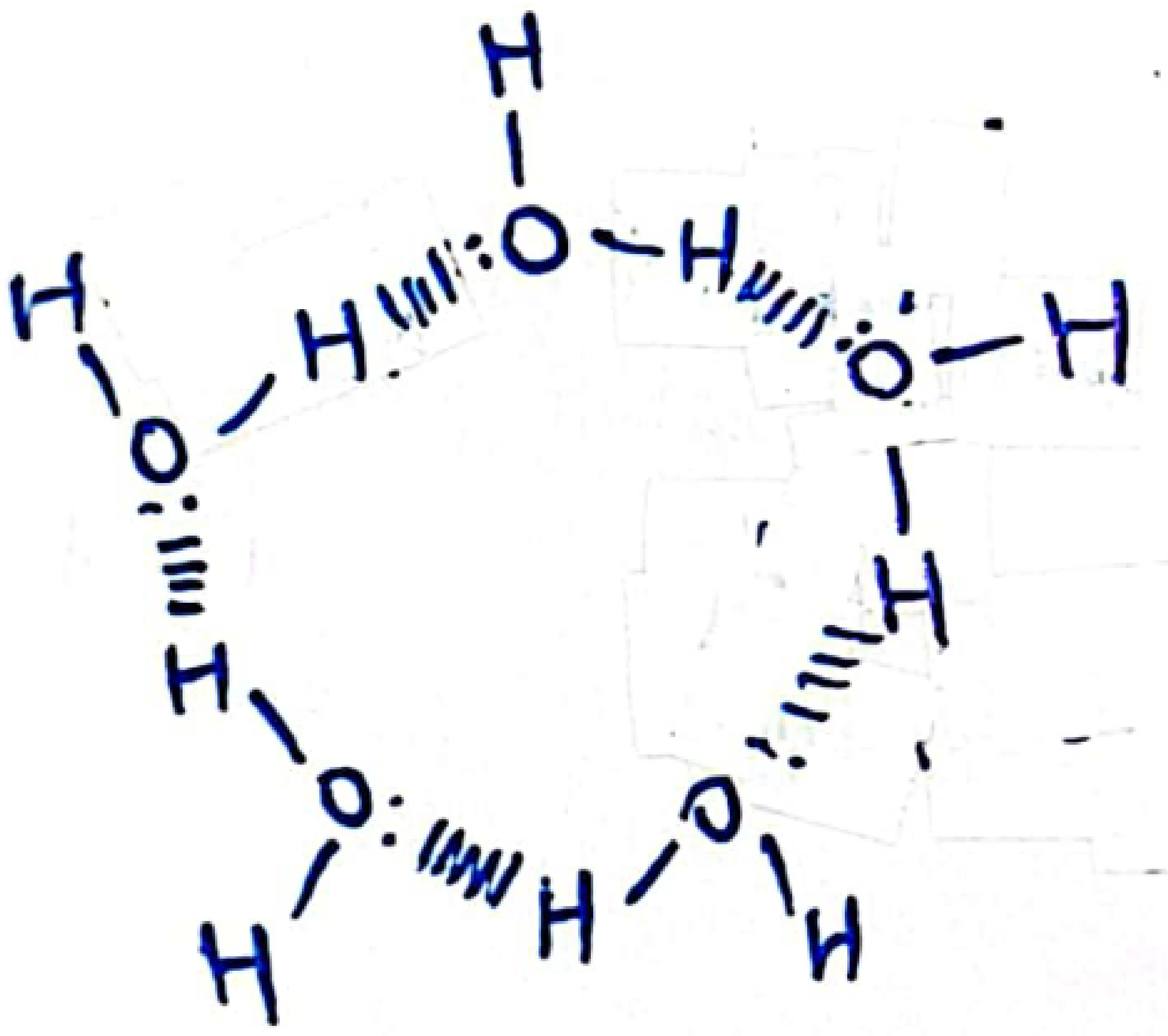
diamond

giant lattice w/ each C atom covalently bonded to 4 other C atoms in a tetrahedral arrangement w/ strong C–C bond



graphite

- giant covalent lattice w/ layers of atoms w/ each C atom covalently bonded to 3 other C atoms in a trigonal planar arrangement
- strong C–C bond, w/ weak id-id FoA btwn layers

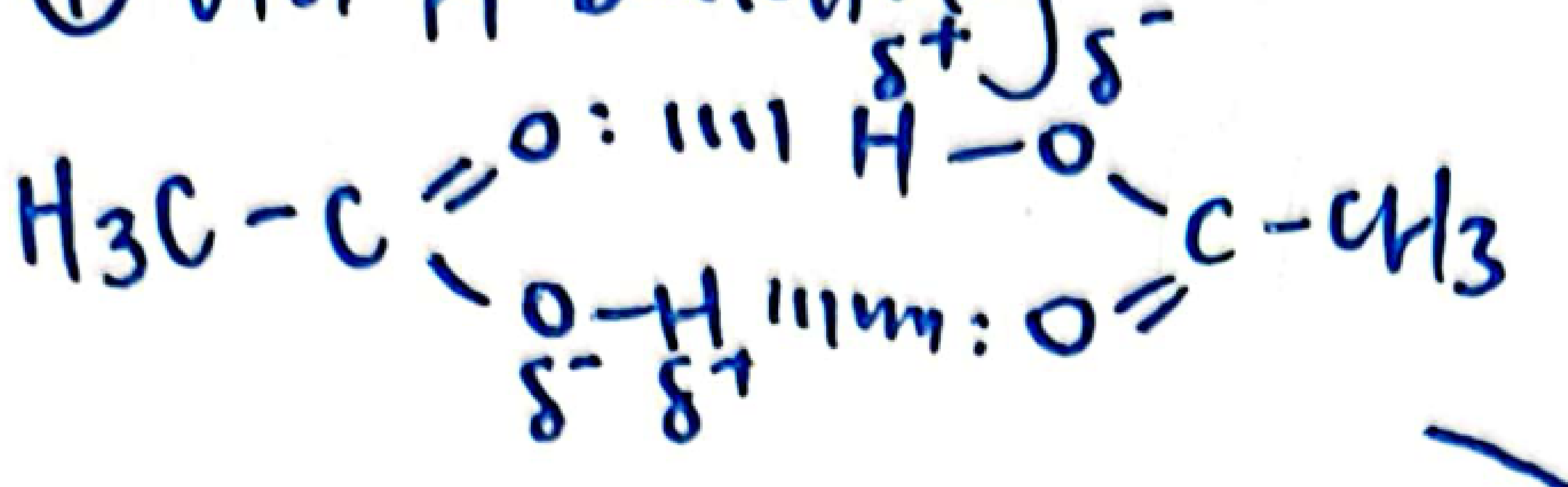


ice

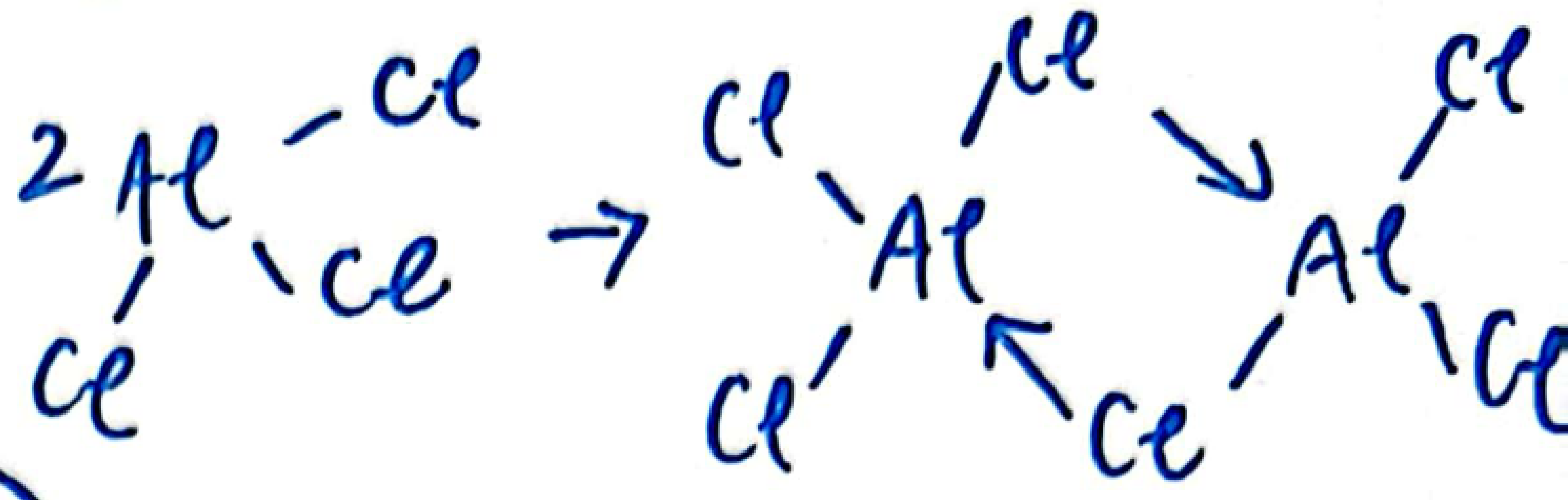
- giant lattice of H_2O molecules, w/ each O atom surrounded tetrahedrally by 4 H atoms (2 via cov. bond, 2 via H bond)
- open lattice str → empty spaces btwn H_2O mol.
  ∴ $H_2O(s)$ less dense than $H_2O(l)$
- strong O-H cov bond, relatively strong H bond btwn H_2O molecules

Dimerisation

① via H bonding



② via dative bond

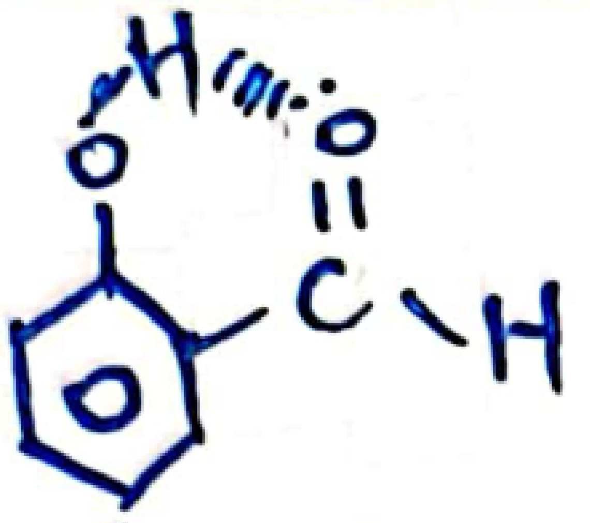


③ via cov bond w/ radicals

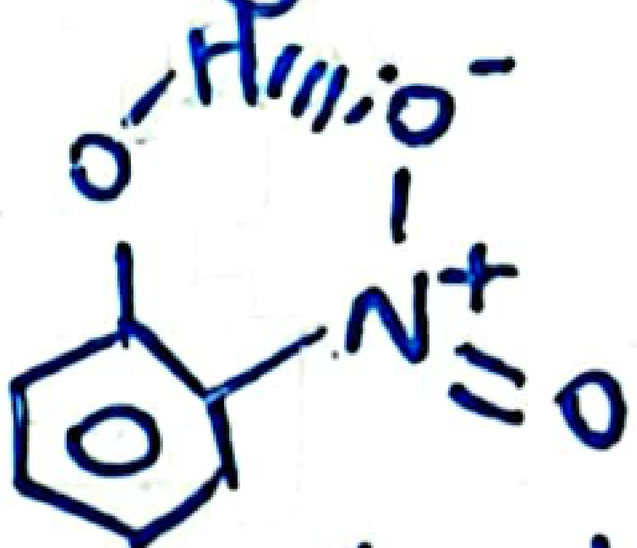
$2\dot{N}O_2 \rightarrow N_2O_4$

CH_3COOH forms H bond w/ another molecule of CH_3COOH, forming dimer w/ Mr 120.0 in non-polar organic solvents

Intramolecular H bonding



2-hydroxybenzaldehyde

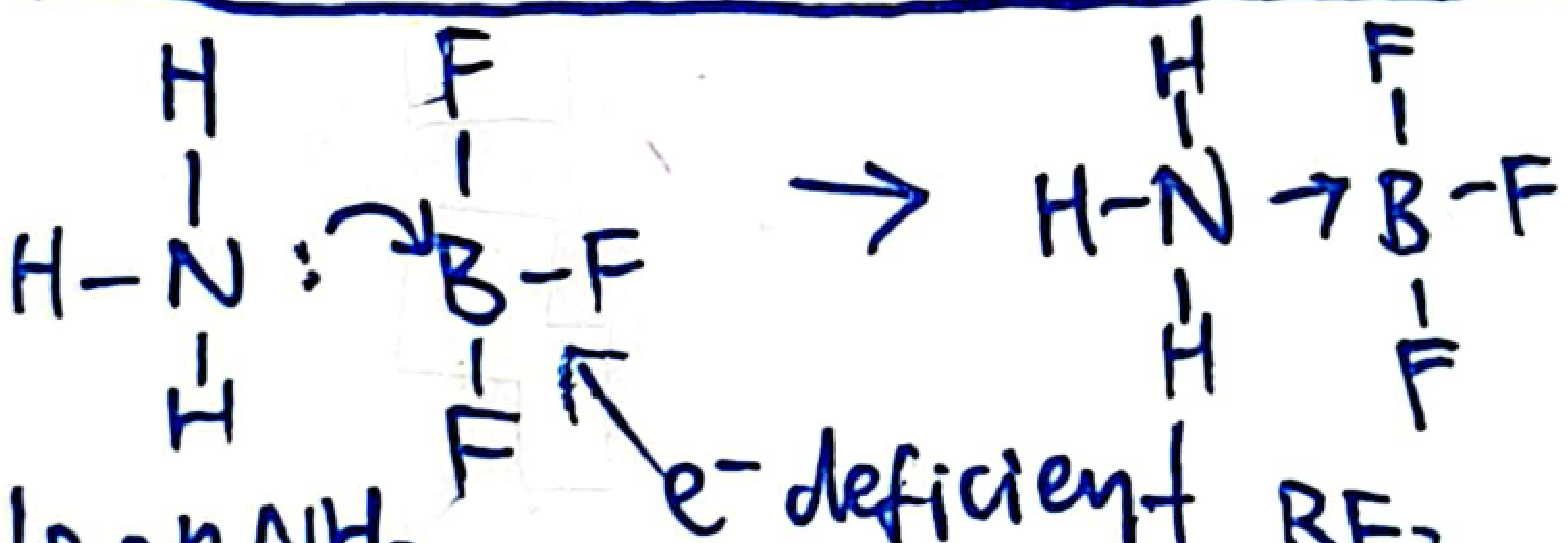


2-nitrophenol

- due to closeproximity of OH & — grps, intramol. H bond can be formed
- intermolecular H bonding less extensive

∴ less energy to overcome less extensive intermol. H bonding.

Dative bond in e⁻ deficient atoms



lp on NH_3 e⁻ deficient BF_3

Chemical Energetics

- $\Delta H^{\ominus}_{rxn}$ is the energy change when molar quantities of reactants as stated in the thermochemical equation react under std. conditions
- ΔH°_{f} is the energy change when one mole of a compound is formed from its constituent elements in their standard states under std. conditions 1 elements ( ) → 1 cpd ( )
- ΔH°_{c} is the energy change when one mole of a substance is completely burnt in excess O_2 under std. conditions 1 cpd + nO_2 → CO_2 + H_2O ($\Delta H_c = -ve$, except N_2 ∵ $N \equiv N$ extremely strong)
- ΔH°_{hyd} is the energy change when one mole of gaseous ions is surrounded by H_2O molecules, forming a solution at infinite dilution under std conditions
  ($\Delta H_{hyd} = -ve$, ∵ of bonds formed btwn gaseous ions and dipoles at H_2O molecules)
  $|\Delta H_{hyd}| \propto \frac{|z|}{r}$ ← factors that affect str of ion-dipole interaction   1 ion(g) + aq → 1 ion (aq)
- ΔH°_{sol} is the energy change when one mole of cpd is dissolved by solvent such that further dilution produces no more heat change, under standard conditions
  $\Delta H_{sol} = \Delta H_{hyd} - LE$ 1 cpd + aq → ion(aq) + ion (aq)
- ΔH°_{n} is the energy change when one mole of H_2O is formed during the neutralisation of an acid and an alkali, under std. conditions   acid + base → salt + 1 H_2O
  ($\Delta H_n = -ve$, $\Delta H_n = -57 \text{ kJ mol}^{-1}$ for SA/SB rxn)
- $\Delta H^{\ominus}_{atm}$ is the energy change when one mole of gaseous atoms is formed from its elements, under standard conditions   element → 1 atom (g)
  ($\Delta H_{atm} = +ve$, bonds are broken, no bonds formed)
- Bond Energy is the energy absorbed to break one mole of covalent bond between two atoms in gaseous state (BE = +ve, covalent bonds broken, no bonds formed)
  1 covalent cpd → 2 atoms
- Ionisation Energy is the energy change when one mole of electrons is removed from one mole of gaseous atoms to form one mole of singly charged gaseous cations
  $X^{n}(g) \rightarrow X^{(n+1)+}(g) + e^-$ ($\Delta H = +ve$, energy absorbed to overcome electrostatic attraction btwn e^- & nucleus to remove e^-)
- Electron Affinity is the energy change when one mole of electrons is added to one mole of gaseous atoms to form one mole of gaseous anions
  $X^{n}(g) + e^- \rightarrow X^{(n-1)-}(g)$ (1st EA $\Delta H = -ve$, 2nd and more EA $\Delta H = +ve$ ∵ of energy to overcome repulsion btwn e^- & anion)
- Lattice Energy is the energy change when one mole of ionic cpd is formed from its constituent gaseous ions at std conditions
  ion(g) + ion(g) → 1 ionic cpd (LE = −ve, ∵ energy evolved when ions come tgt to form ionic bonds)
  $|LE| \propto \frac{z^+ z^-}{r_+ + r_-}$

Discrepancies for calculated ΔH

- ΔH determined experimentally
  - ↳ heat gain from surroundings (endo) / heat loss to surroundings (exo)
  - ↳ rxn does not occur rapidly enough for max T to be achieved before rxn mixture begins to return to room temp
- ΔH determined by BE
  - ↳ not in gaseous state + data booklet values is avg value
- LE
  - ↳ degree of covalent character

ΔH_c → insufficient O_2 → incomplete combustion

- Hess' Law: enthalpy change for a chemical rxn is the same regardless of the route the reaction takes, provided the initial states of the reactants and the final states of the products are the same.

- Entropy is the measure of disorder in a system

Factors affecting entropy

- Δ temperature
  - ↳ ↑ temp → system possess more KE → more ways to distribute energy packets among particles → ↑ disorder in system → ↑ entropy
- mixing of particles
  - ↳ gas: more ways of arranging the gaseous particles → ↑ disorder → ↑ entropy → $\Delta S = +ve$
  - ↳ dissolving: (solid) broken from orderly arrangement in solid lattice → can move around freely when dissolved → more ways of arranging particles → ↑ disorder → ↑ entropy → $\Delta S = +ve$
- Δ phase (eg solid → liquid) → more ways to arrange particles → ↑ disorder → ↑ entropy → $\Delta S = +ve$
- no. of particles (gases)
  - ↳ ↑ no. of particles → more ways to arrange molecules → ↑ disorder → ↑ entropy → $\Delta S = +ve$
- Δ volume of gas
  - ↳ ↑ volume → more ways to arrange molecules → ↑ disorder → ↑ entropy → $\Delta S = +ve$

Spontaneity ($\Delta G = \Delta H - T\Delta S$)

$\Delta G < 0$: spontaneous
$\Delta G > 0$: non-spontaneous

} Assumption :
- ΔH & ΔS independent of temp.
- valid so long no Δ phase

Effect of ΔT on ΔG

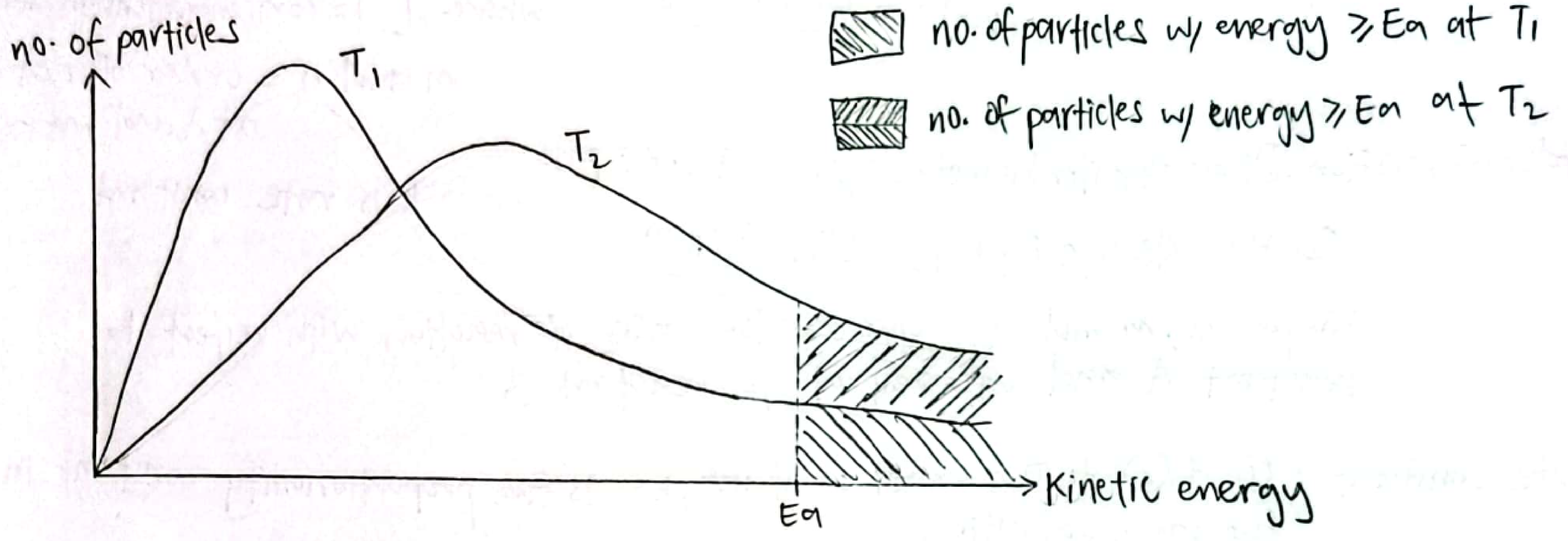
1. gives signs of ΔH, ΔS and −TΔS
2. $\Delta G = \Delta H - T\Delta S$
3. at high temp., $|T\Delta S| > |\Delta H|$ / $-T\Delta S$ becomes more +ve/−ve
4. ∴ ΔG becomes +ve/−ve at high temp.
5. rxn is more/ less spontaneous at high temp

$$\Delta G^\circ = -RT \ln K$$

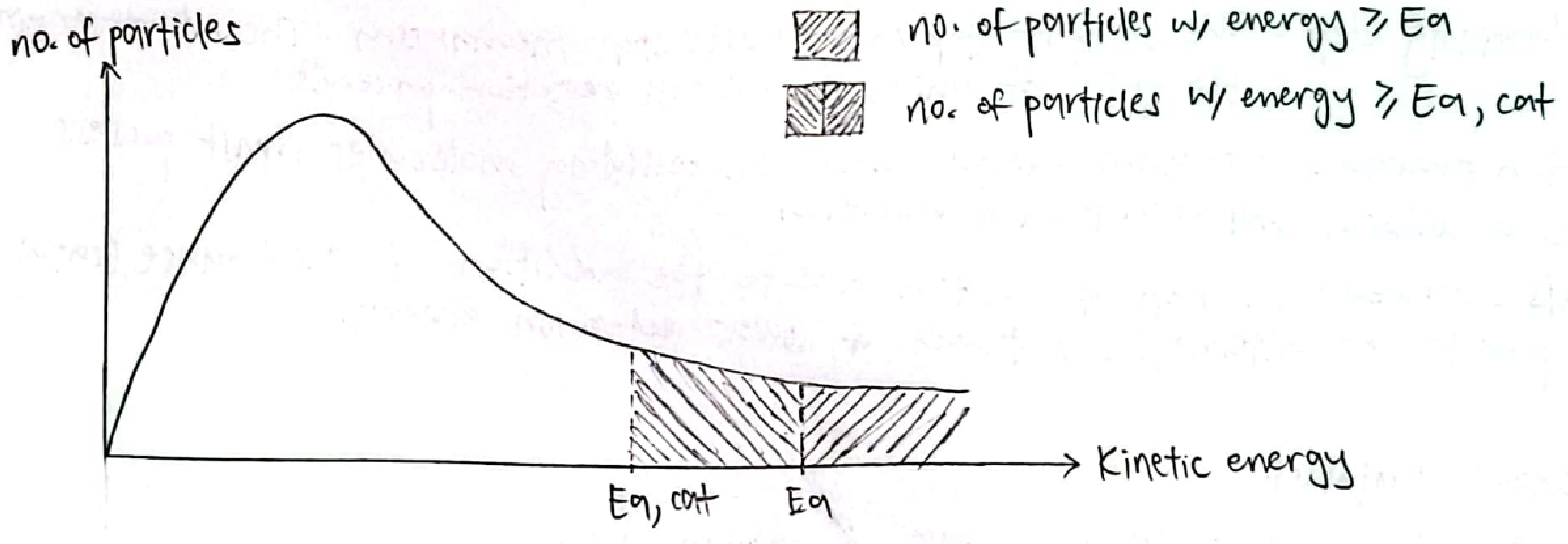
Reaction Kinetics

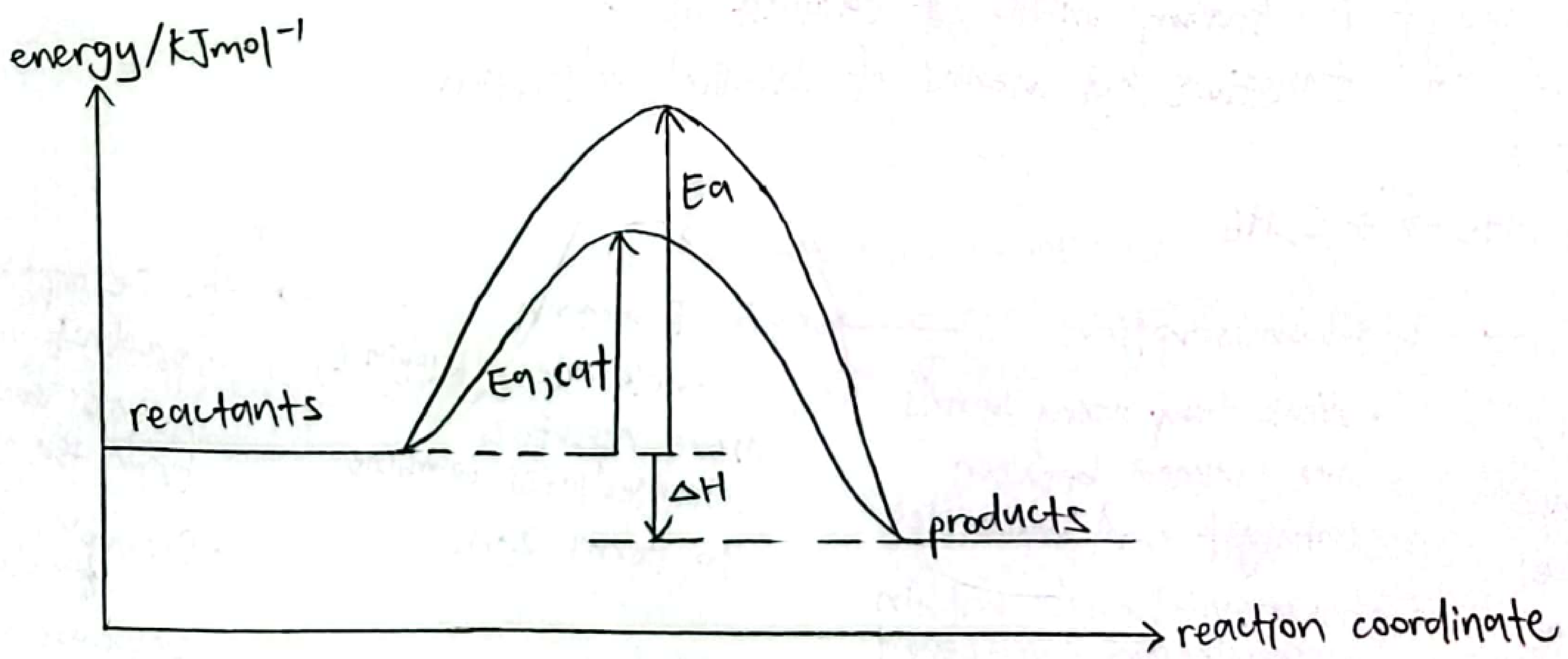
Factors affecting rate

- Δ[reactant]
  - ↑[ ] → reactant closer to one another → freq of effective collisions ↑ → rate ↑
  - Exceptions : - rxn is zero order wrt rxt
    - catalyst already working at max capacity
- ΔTemp
  - temp ↑ → avg KE ↑ → more particles have energy ≥ Ea → can be seen by the flattening of Maxwell-Boltzmann curve → · freq of effective collisions ↑ → rate const ↑ → rate ↑



- Catalyst
  - catalyst added → provides alternative pathway with lower Ea → ↑ particles w/ energy ≥ Ea → frequency of effective collisions ↑ → rate const. k ↑ → rate ↑





Factors affecting rate constant 'k'

- Arrhenius equation : $K = Ae^{-\frac{Ea}{RT}}$ Ea → activation energy
  T → temperature
- Temperature ↑ → $e^{-\frac{Ea}{RT}}$ ↑ → k↑
- catalyst → Ea↓ → $e^{-\frac{Ea}{RT}}$ ↑ → k↑

- Rate of reaction : change in concentrations of reactants or products per unit time   units : $mol dm^{-3} time^{-1}$

- Rate equation : For a general rxn: aA + bB → products
  Rate equation is,
  $$rate = k[A]^m[B]^n$$
  where [ ] = concentrations in $mol dm^{-3}$
  m and n = order of rxn wrt rxt A and wrt rxt B
  k is rate constant

- order of reaction: ① For the stoichiometric eqn : aA + bB → pdts
  ② the rate eqn is, $rate = k[A]^m[B]^n$
  The powers, m and n, represent the orders of reaction with respect to reactant A and with respect to reactant B

- rate constant : ① + ② + The rate constant, k, is the proportionality constant in the rate equation

- Half-life, $t_{\frac{1}{2}}$: time taken for the conc. of the reactant to fall to half of its original value $t_{1/2} = \frac{\ln 2}{k}$

- Rate determining step : The slowest step in a multi-step mechanism. The rds determines the speed at which the overall reaction proceeds

- Activation energy : minimum energy which the colliding molecules must possess before a collision will result in a reaction

- Catalysis : increase in rate of reaction due to the addition of a substance (catalyst) which provides an alternative pathway of lower activation energy.

Heterogenous Catalysis

- reactant and catalyst in different phase + catalysis defⁿ
- usually a solid and functions by means of chemical adsorption

MODE OF ACTION

Eg: $C_2H_4 + H_2 \xrightarrow{Ni} C_2H_6$

1. Diffusion →
   - Reactant molecules, C_2H_4 & H_2, diffuse towards the nickel catalyst surface
   - Adsorb onto the surface of catalyst

2. Chemisorption →
   - Weak temporary bonds are formed between catalyst and molecules
   - covalent bonds within molecules weaken, lowering Ea

3. Reaction →
   molecules dissociate; more reactive intermediates combine to form C_2H_6

4. Desorption
   - product molecules break away from the surface
   - Desorp away from ~~g~~& diffuse catalyst surface

Homogenous Catalysis

- reactant and catalysts are in same phase + catalysis def^{n}

MODE OF ACTION

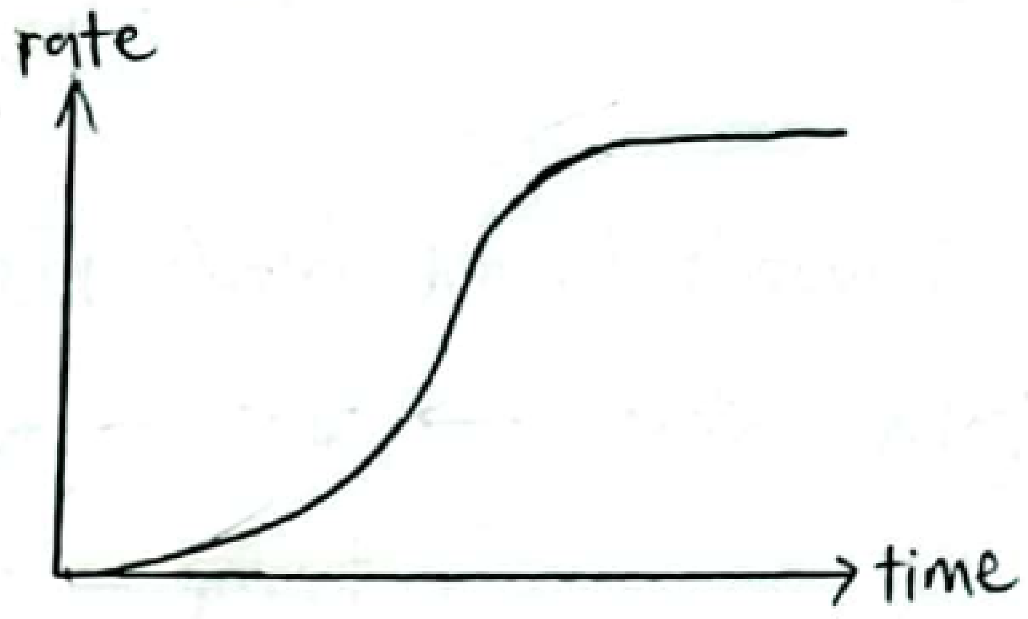
- A homogenous catalyst first forms an intermediate compound w/ one of the reactants which then reacts with the other reactant to form the final products & regenerate back the catalyst

i.e. 1. A + catalyst → intermediate   Formation of intermediate

2. intermediate + B → product + catalyst   Regeneration of catalyst

overall : A + B → product

Autocatalysis → product acts as catalyst



Examples for catalysis

- Heterogenous
  - ↳ Haber process
    - Fe(s) catalyst
    - Eqn: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$
    - mode of action
  - ↳ Removal of oxides of nitrogen in exhaust gases from car engines
    - Platinum catalyst
    - Eqns: $CO + NO \rightarrow \frac{1}{2}N_2 + CO_2$

      $2CO + NO_2 \rightarrow \frac{1}{2}N_2 + 2CO_2$

      $NO_2 \rightarrow \frac{1}{2}N_2 + O_2$

      $NO \rightarrow \frac{1}{2}N_2 + \frac{1}{2}O_2$
    - mode of action
- Homogenous
  - ↳ Catalytic role of atmospheric oxides of nitrogen in the oxidation of atmospheric SO_2
    - NO_2 catalyst
    - Eqns: Step 1 : $SO_2 + NO_2 \rightarrow SO_3 + NO$

      Step 2: $NO + \frac{1}{2}O_2 \rightarrow NO_2$

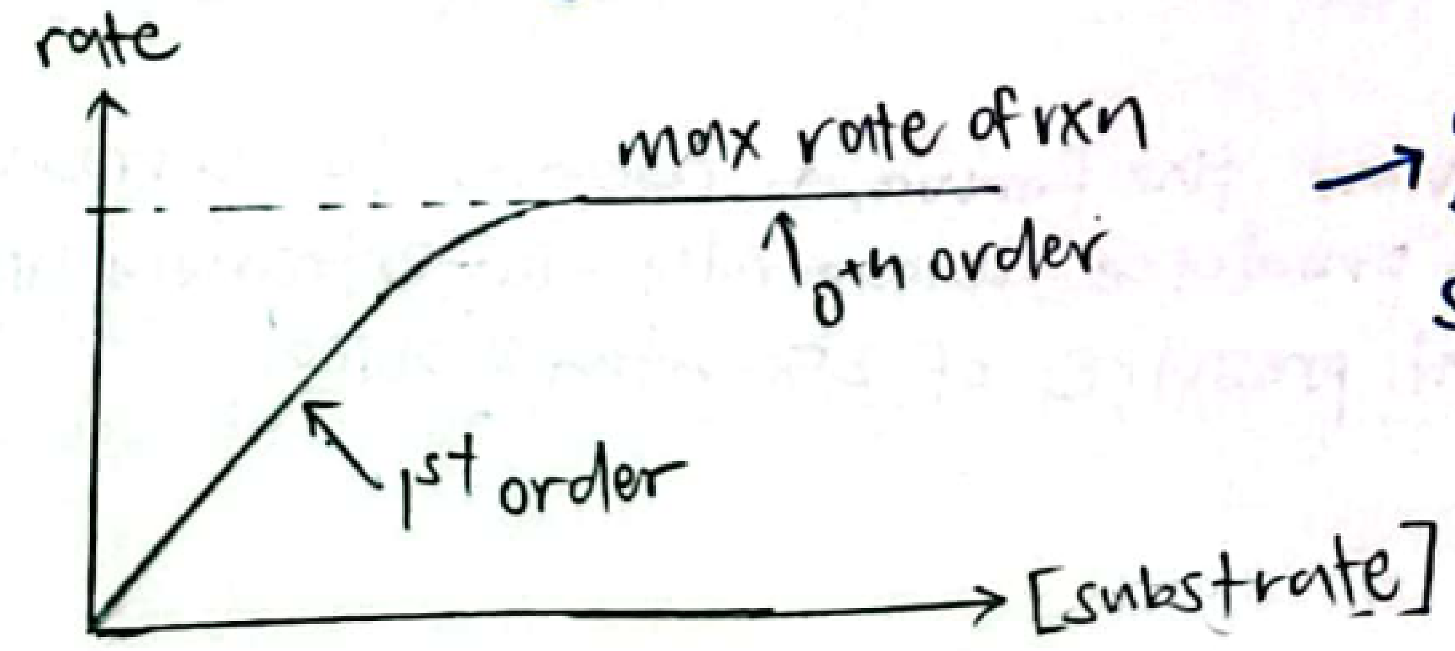
      Overall : $SO_2(g) + \frac{1}{2}O_2(g) \rightarrow SO_3(g)$
  - ↳ Catalytic role of Fe^{2+} in $I^- / S_2O_8^{2-}$ reaction
    - Fe^{2+} catalyst
    - Eqns : Step 1 : $S_2O_8^{2-} + 2Fe^{2+} \rightarrow 2SO_4^{2-} + 2Fe^{3+}$

      Step 2 : $2Fe^{3+} + 2I^- \rightarrow 2Fe^{2+} + I_2$

      Overall : $S_2O_8^{2-}(aq) + 2I^-(aq) \rightarrow 2SO_4^{2-}(aq) + I_2(aq)$

Enzymes

- biological catalysts that catalyse biochemical reactions
- binds substrate at active site, in correct orientation → binding weakens bonds within substrate molecule, lowering Ea of reaction
- temp. and pH sensitive



→ at high substrate conc., the enzyme active sites are FULLY SATURATED ∴ max rate

... over

Chem Equilibria

- Reversible reaction: a reaction that proceeds in both the forward and backward direction. The reaction contains quantities of all reaction species in the system
- Dynamic equilibrium: In a reversible system, dynamic equilibrium occurs when the rates of the forward and backwards reactions are the same. There is no net change in the concentration of the reactants and products
- Le Chatelier's Principle states that when a stress is applied to a reversible system at equilibrium, the position of equilibrium will shift to minimise the stress

LCP Explanation

1. By Le Chatelier's Principle, (change in condition)
   ↳ ↑/↓ temp / pressure / conc.
2. The forwards / backwards reaction is favoured to
3. Explain why → conc − to produce / use up ___
   temp − exo to produce heat
   endo to absorb heat
   pressure − to ↑/↓ no. of gaseous particles
4. causing POE to shift left/ right
5. Describe any observations (if qn ask for it)

ΔG° & POE

$\boxed{\Delta G^\ominus > 0}$	$\boxed{\Delta G^\ominus = 0}$	$\boxed{\Delta G^\ominus < 0}$
− [reactants] >> [products]	− [reactants] = [products]	− [reactants] << [products]
− K << 1	− K = 1	− K >> 1
− POE lies very much to the left, most of reactants left unreacted		− POE lies very much to the right, reaction is complete

Haber Process

$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$ $\Delta H = -92.4\ kJ\ mol^{-1}$

Conditions: Temp → 500°C
Pressure → 250 atm
Catalyst → finely divided iron catalyst

Reason for conditions

- Temp
  ↳ By LCP, ↓ temp will favour forwards exothermic rxn to produce more heat, POE shift right to produce more NH_3. But at low temp, rate of rxn is slow & system takes long time to reach equilibrium, ∴ optimal temp of 500°C and catalyst used to ensure both yield and rate is high
- Pressure
  ↳ By LCP, ↑ pressure will favour the forward reaction to decrease the total amt of gases, POE shifts to the right to produce more NH_3. But extremely high pressures will involve ↑cost and equipment ∴ optimal pressure of 250 atm is used

Ionic Equilibrium

- Arrhenius acid / base forms H^+/OH^- ions in aqueous solution
- Lewis acid / base is an electron pair acceptor / donor
- Bronsted-Lowry acid / base is a proton donor / acceptor
- Weak acids / bases partially dissociates in water to produce a low concentration of H^+/OH^- ions
- Strong acids / bases completely dissociates in water to produce a high concentration of H^+/OH^- ions
- $pH = -\lg[H_3O^+]$ $pOH = -\lg[OH^-]$ $pH + pOH = 14.0$ $K_w = K_a \times K_b$ $pK_w = pK_a + pK_b$

Buffer

Buffer resists large changes in pH when small amounts of acid or base is added.

- $pH = pK_a + \lg\frac{[salt]}{[acid]}$ $pOH = pK_b + \lg\frac{[salt]}{[base]}$

Acidic / Basic buffer : mixture of weak acid / base and conjugate base / acid

- When small amount of acid is added : Most of the added H_3O^+ reacts with the large reservoir of A^-/B, keeping pH fairly constant
  - ↳ $H_3O^+ + A^- \rightarrow HA + H_2O$
  - ↳ $H_3O^+ + B \rightarrow BH^+ + H_2O$
- When small amount of base is added : Most of the added OH^- reacts with the large reservoir of HA/BH^+, keeping pH fairly constant
  - ↳ $OH^- + HA \rightarrow A^- + H_2O$
  - ↳ $OH^- + BH^+ \rightarrow B + H_2O$

maximum buffering capacity (MBC)
- occurs when $[salt] = [acid]/[base]$
- $pH = pK_a$ / $pOH = pK_b$

- H_2CO_3/HCO_3^- buffer system in blood

$H_3O^+ + HCO_3^- \rightarrow H_2CO_3 + H_2O$

$OH^- + H_2CO_3 \rightarrow HCO_3^- + H_2O$

Choosing indicators → sharp colour change at equivalence point on adding single drop of titrant

Explanation: the working range of indicator lies within the sharp pH change at equivalence point

Methyl orange 3-4 $\overset{basic}{\text{yellow}} \rightarrow \overset{\star\ end point}{\text{orange}} \rightarrow \overset{acidic}{\text{red}}$

Phenolphthalein 8-10 $\overset{acidic}{\text{colourless}} \rightarrow \overset{\star\ end point}{\text{pink}} \rightarrow \overset{basic}{\text{red}}$

Salt hydrolysis (to answer whether (salt) forms acidic / basic / neutral solution)

- Acids (eg NH_4^+, CH_3CO_2H)

$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$

NH_4^+ hydrolyses in water to produce H_3O^+ ions, forming an acidic solution

- Bases (eg $CH_3CO_2^-$, NH_3)

$CH_3CO_2^- + H_2O \rightleftharpoons CH_3CO_2H + OH^-$

$CH_3CO_2^-$ hydrolyses in water to produce OH^- ions, forming a basic solution

- Solubility is the maximum amount of substance that can dissolve in a solvent at a given temperature to form a saturated solution

$$\text{solubility} = \frac{\text{max amt that can dissolve}}{\text{volume of solvent}}$$

- Solubility product, K_{sp}, of a sparingly soluble salt is the product of the equilibrium concentrations of the ions in a saturated solution, each raised to the power of its coefficient in the equilibrium equation

For a sparingly soluble salt A_xB_y,

$K_{sp} = [A^{y+}]^x[B^{x-}]^y$

}
- varies w/ temp
- ↓ K_{sp}, ↓ solubility

Factors Affecting Solubility

- Common Ion Effect

eg. decreased solubility of PbI_2 in KI

$PbI_2 \rightleftharpoons Pb^{2+} + 2I^- \quad -(1)$

$KI \rightarrow K^+ + I^-$

Presence of common ion I^- increases $[I^-]$ due to KI. By LCP, POE in (1) shifts left, reducing $[I^-]$ ∴ solubility of PbI_2 reduced

- Formation of Complex Ion

eg solubility of Ag in NH_3

$AgCl \rightleftharpoons Ag^+ + Cl^- \quad -(1)$

when NH_3 added, Ag^+ is removed to form soluble $[Ag(NH_3)_2]^+$ complex ion, causing $[Ag^+]$ in (1) to decrease, POE of (1) shifts right and ppt dissolves.

when IP of AgCl < K_{sp} of AgCl, AgCl dissolves completely

$Ag^+ + 2NH_3 \rightleftharpoons [Ag(NH_3)_2]^+$

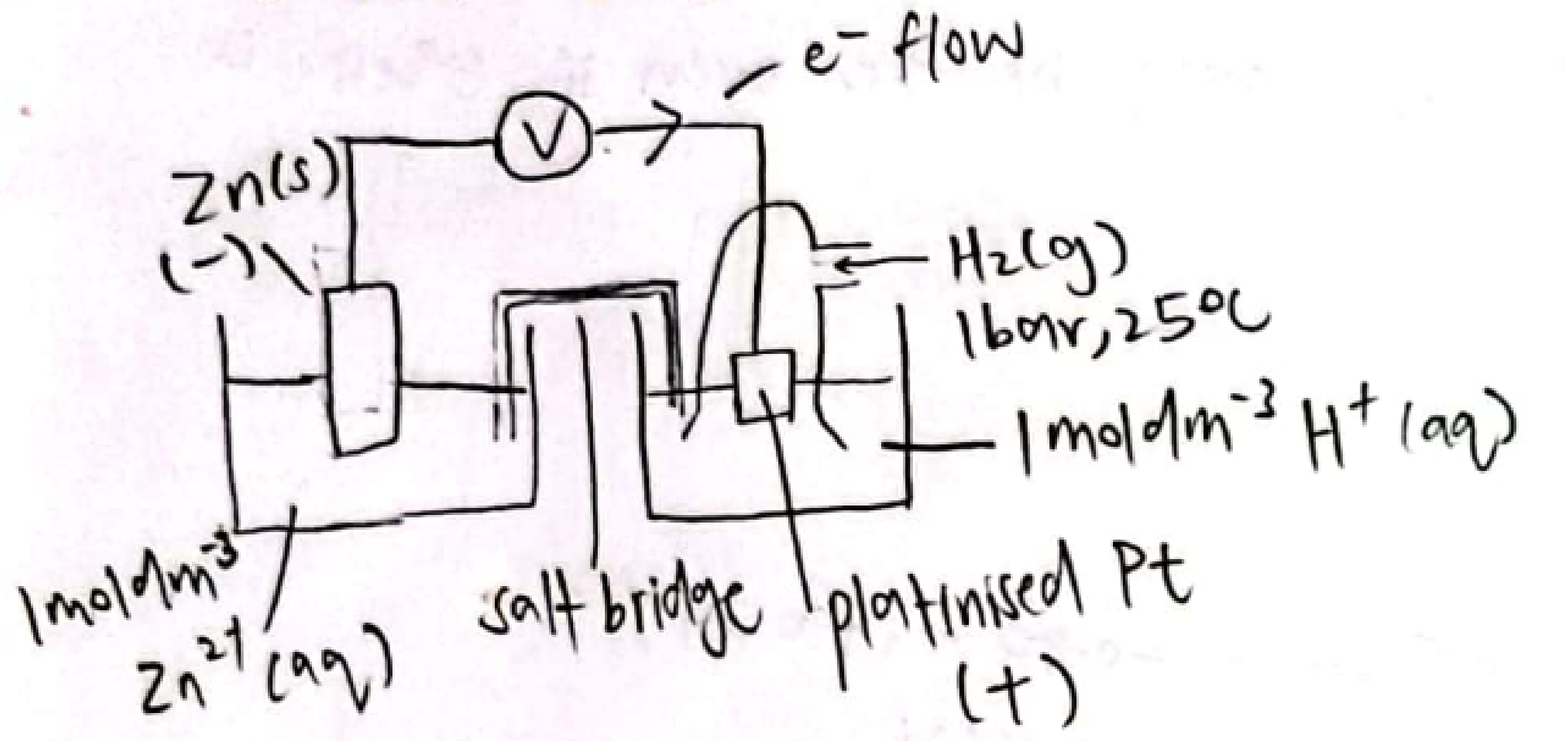
Electrochemistry

- Standard Electrode Potential, E° of half cell is the maximum potential difference (pd) between electrode potentials of the half cell and standard hydrogen electrode (SHE) measured under standard conditions
  ↳ 25°C, 1 bar, [ions] = 1 mol dm^{-3}
- Standard Cell Potential, E°cell is the maximum potential difference between two electrodes of a galvanic cell measured under standard conditions

Half cells

Metal / Ion

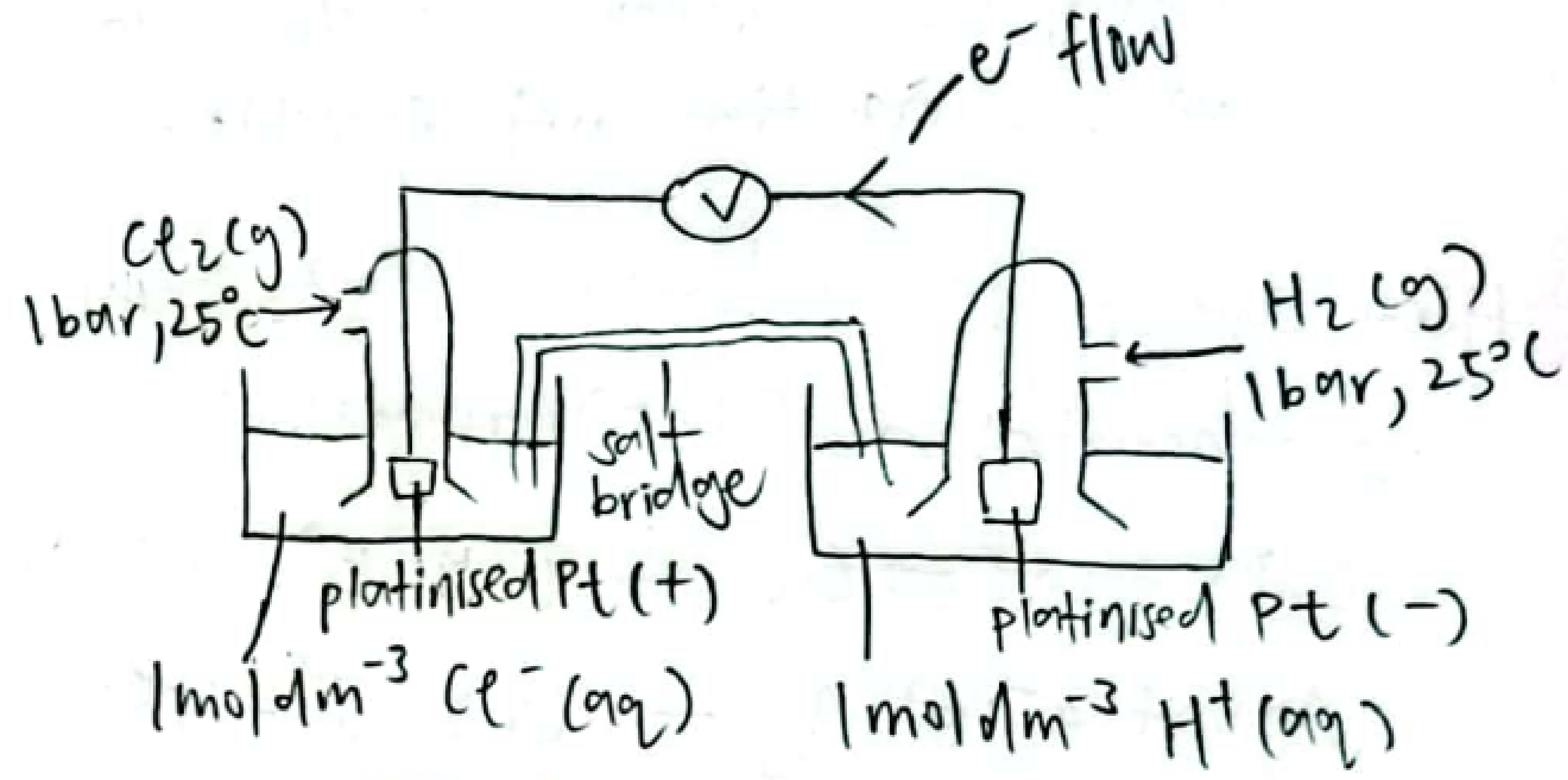
$E^\circ(Zn^{2+}/Zn) = -0.76V$



∴ Zn^{2+}/Zn undergo oxidation

Gas / Ion

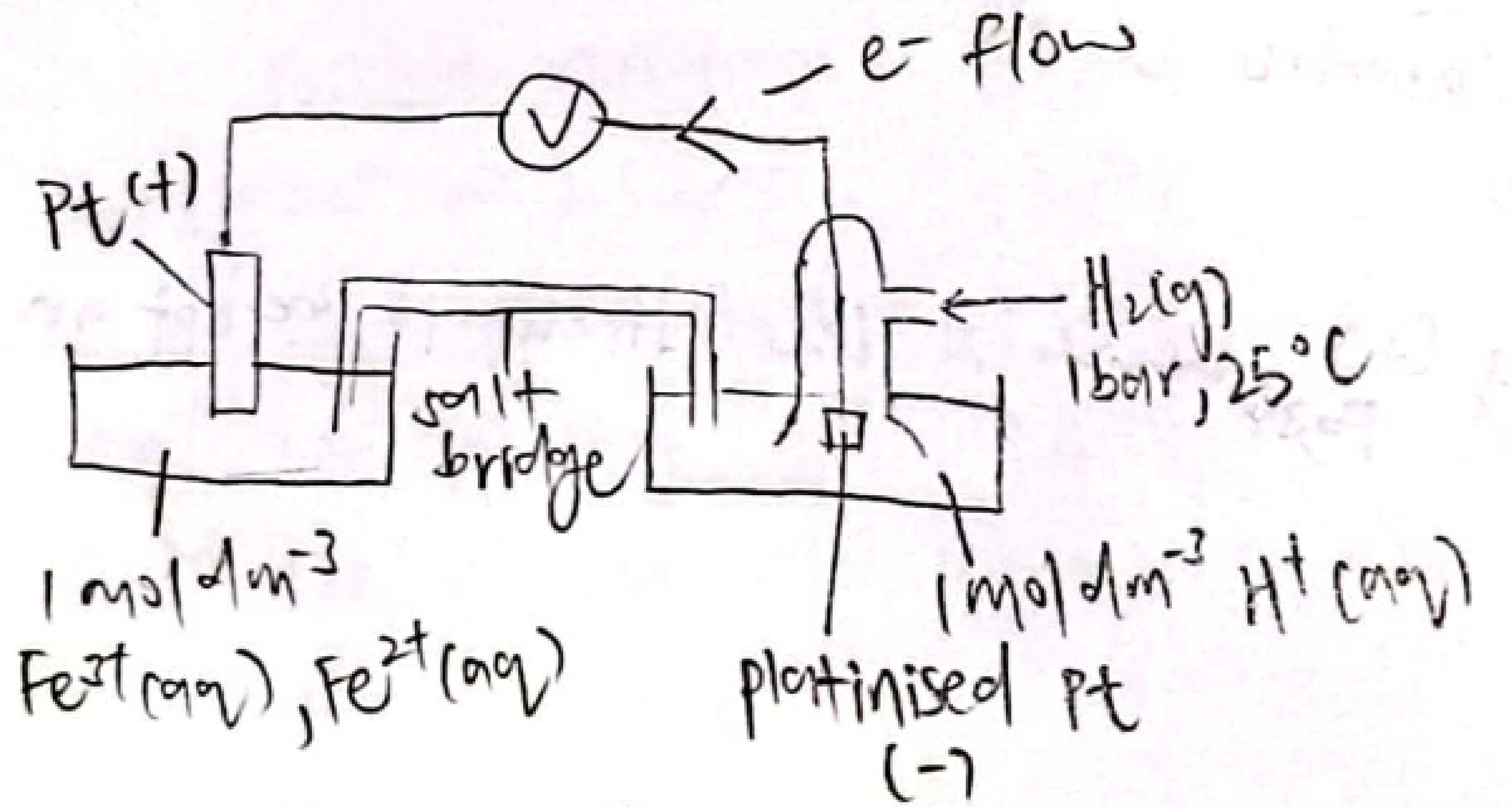
$E^\circ(Cl_2/Cl^-) = +1.36V$



∴ Cl_2/Cl^- undergo reduction

2 Ions

$E^\circ(Fe^{3+}/Fe^{2+}) = +0.77V$



∴ Fe^{3+}/Fe^{2+} undergo reduction

Advantages of batteries in electric cars
- electric motor 90% efficient, petrol only 25% efficient
- delivers a large current for a short time
- provides a high voltage
- rechargeable

Advantages of a H_2/O_2 fuel cell
- high power efficiency (up to 70% efficiency)
- no pollution ∵ no need burn fuel

Strength of OA/ RA
- the more positive / negative the value of $E^\circ(X/Y)$
- the more reduction / oxidation of ___ is favoured
- the stronger the oxidising / reducing agent

~~M stronger oxidising str. Atg~~
~~X_2 stronger reducing str. Atg~~

Preferential reaction for electrolysis
- since $E^\circ(O_2/H_2O) = +1.23V$ is less than $E^\circ(Cl_2/Cl^-) = +1.36$,
  H_2O is preferentially oxidised at anode / Cl_2 preferentially reduced at cathode to form O_2 gas / Cl^- ions

Intro to Organic Chem

Cis-trans isomerism

- occurs when there is restricted rotation about
  - ↳ C=C bond in an aliphatic alkene or
  - ↳ C-C bond in a small rigid ring
- and when there are two different groups on each carbon of the double bond

Enantiomers – pair of chiral compounds that are isomers
– rotate plane polarised light

stereoisomers ✓

Optically active molecules have:

1. a chiral carbon (4 diff atoms bonded to it)
2. non-superimposable mirror images
3. no plane, axis or centre of symmetry

Constitutional isomerism

- compounds having same molecular formula but have different structural formula

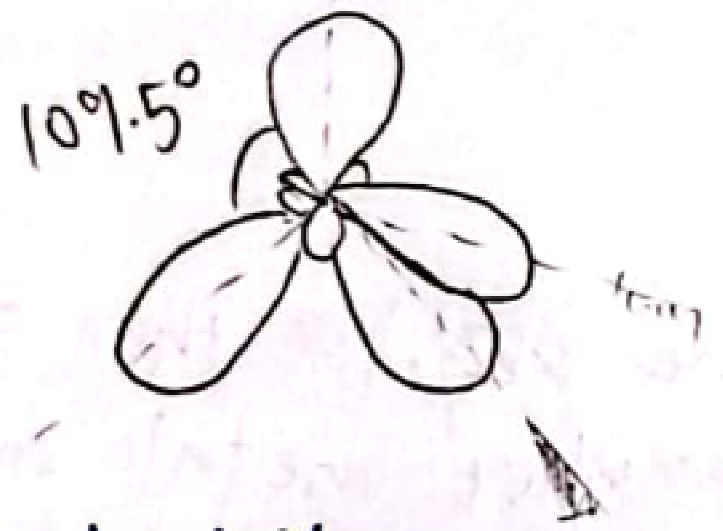
↓

Stereoisomerism

- compounds with same molecula formula but with different spatial formula

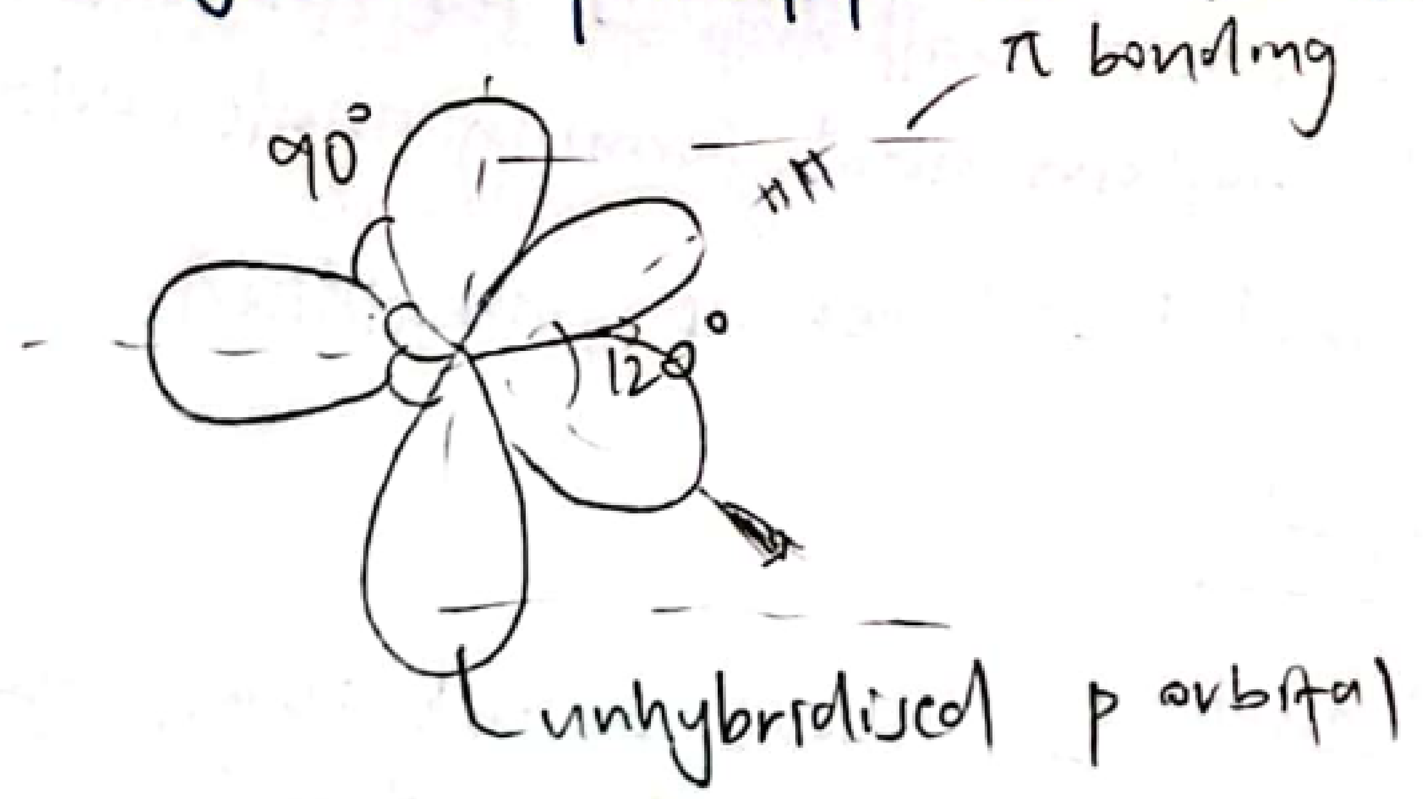
sp^3 hybridisation (ethane)

- formed through combination of one s and three p orbitals
- tetrahedral shape



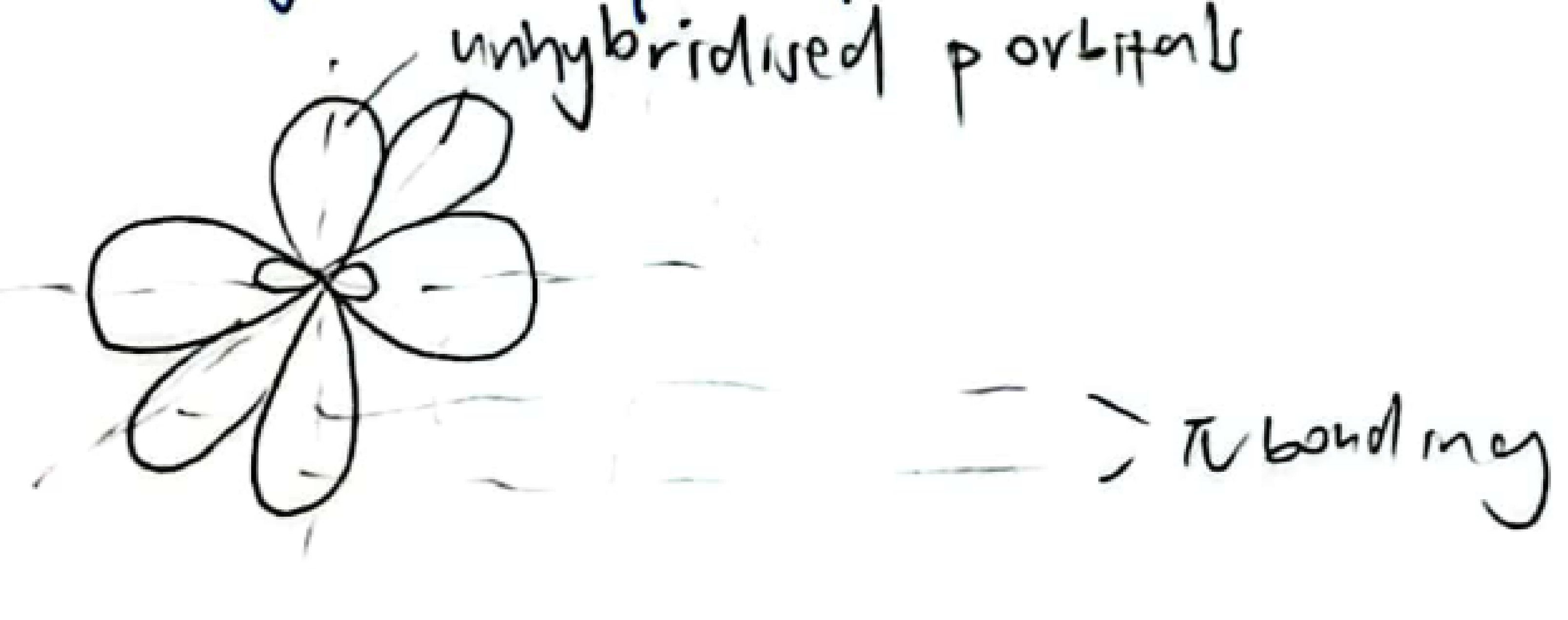
sp^2 hybridisation (ethene)

- formed through combination of one s and two p orbitals
- unhybridised p orbital used for π bonding



sp hybridisation (ethyne)

- formed through combination of one s and one p orbital
- two unhybridised p orbitals used for π bonding



Alkanes

Homolytic Fission - process where a covalent bond is broken, forming two radicals

Heterolytic fission - breaking of covalent bond, and more electronegative atom takes both bonding electrons to form an anion, leaving a cation

Why Alkanes are unreactive:

1. Non polar bond
   ↳ C-H bond is non-polar → alkane molecules do not attract reactive species
2. Bond strength
   ↳ C-C and C-H bond relatively strong

Undergo FRS because:

· alkanes saturated → does not undergoes addition rxn ~~to form a single product~~

Brief exposure to UV light / Why X· is a homogenous catalyst

· UV light produces X· radicals that initiate chain reactions of the propagation step to produce more radicals for reaction to sustain

Alkenes

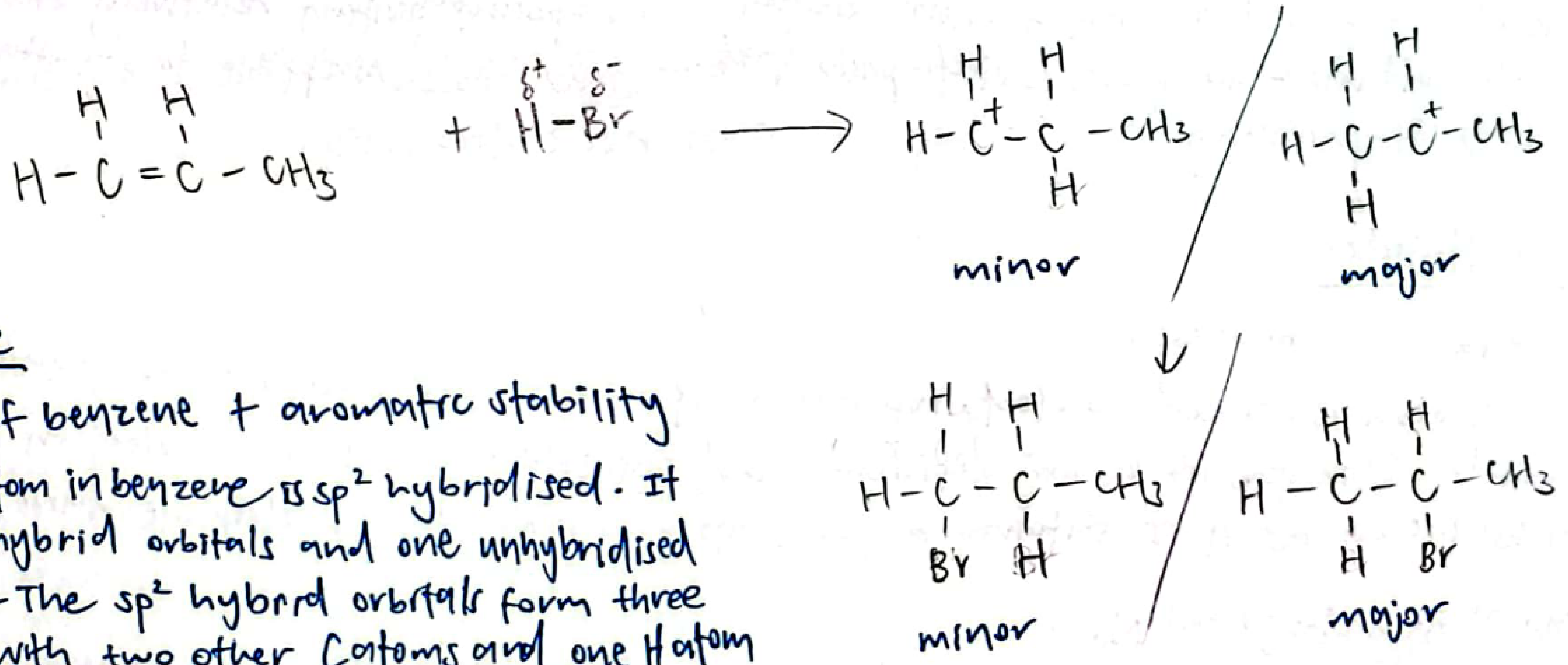
Undergo EA because:

1. e^- rich C=C bond attracts e^- deficient electrophiles
2. C=C bond in alkene is unsaturated ∴ a small molecule can be added across to form a single pdt

Markonikov's Rule → presence of more e^- donating alkyl grp → ↑ e^- density carbocation centre → disperse +ve charge, stabilising carbocation

- major product of addition reaction is formed from the most stable carbocation i.e most substituted C atom in C=C

eg

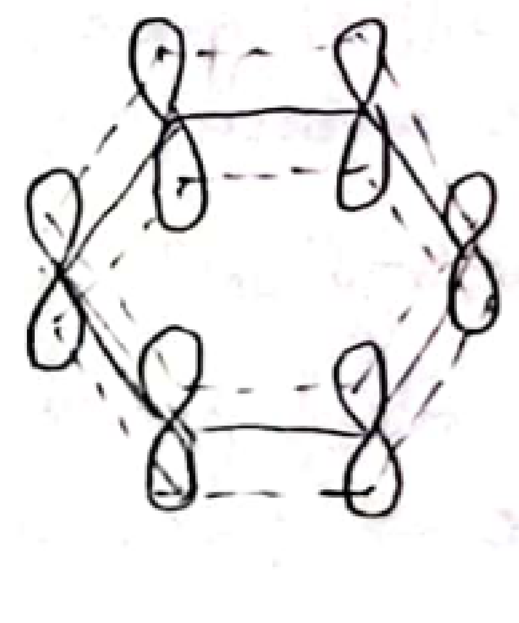


Benzene

Structure of benzene + aromatic stability

· each C atom in benzene is sp^2 hybridised. It has 3 sp^2 hybrid orbitals and one unhybridised p orbital - The sp^2 hybrid orbitals form three σ bonds with two other C atoms and one H atom via head on overlap on the same plane.

- The 6 p orbitals on the C atoms are close enough to one another for side on overlap to form a ring of delocalised π electrons above & below the plane, giving rise to extra stability in benzene



Benzene undergo ES because:

- delocalised π electron cloud in benzene has a large e^- density to attract e^- deficient electrophiles, making it susceptible to electrophilic attack
- undergo substitution to maintain the ring of delocalised π e^- cloud that provides aromatic stability

Role of $AlCl_3$ in ES:

- acts as a lewis acid catalyst by accepting an electron pair from Cl to produce strong electrophile Cl^+

Role of H_2SO_4 in ES:

- acts as a bronsted-lowry acid catalyst, donating a proton to HNO_3 to produce the stronger electrophile NO_2^+

Why $AlCl_3$ anhydrous:

$AlCl_3 + 6H_2O \rightarrow [Al(H_2O)_6]^{3+} + 3Cl^-$

$[Al(H_2O)_6]^{3+} + H_2O \rightleftharpoons [Al(H_2O)_5(OH^-)]^{2+} + H_3O^+$

- $AlCl_3$ undergoes hydration in presence of water, followed by hydrolysis to form an acidic solution. ∴ $AlCl_3$ cannot act as a lewis acid as there is no more energetically accessible and vacant orbitals in Al to accept a lone pair of e^- from Cl

Benzene vs Alkene reactivity

- Benzene reacts less readily w/ electrophiles than alkenes
- ∵ benzene is less e^- rich due to the delocalisation of π e^- cloud

} think stronger X^+ electrophile needed vs X_2 in alkene

Effect of substituents on reactivity & directing effect

- Activating groups → make aromatic ring more reactive than benzene
- e^- donating groups ↑ e^- density in benzene ring at the 2- and 4- positions
- makes 2- and 4- position more electron rich and more susceptible to electrophilic attack
  ∴ 2, 4- directing

- Deactivating groups → make aromatic ring less reactive than benzene
- e^- withdrawing group ↓ e^- density in benzene ring from the 2- and 4- positions
- makes 3-position relatively richer in e^- and more susceptible to electrophilic attack
  ∴ 3-directing

Alkyl Halides

Undergo NS because:

- RX C atom is bonded to electronegative X atom, ∴ C-X bond is polar, causing alkyl halide C centre to be e^- deficient and susceptible to attack by electron rich nucleophiles
- substitution as it is saturated, ∴ X atom is replaced by nucleophile during rxn

Why 3° RX undergo SN1 NS:

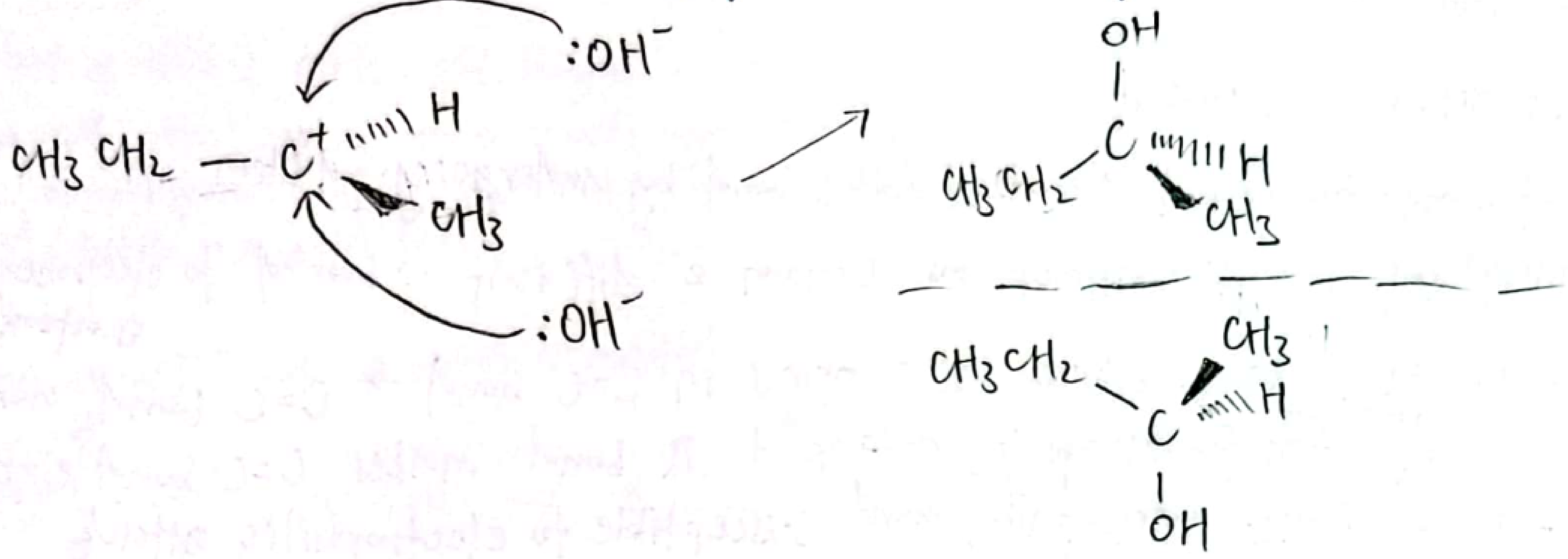
- e^- donating alkyl groups ↑ e^- density of carbocation centre → disperse the positive charge on the carbocation → ∴ stabilising the carbocation

Why 1° RX undergo SN2 NS:

- lesser bulky alkyl groups that crowd the backside of the e^- deficient alkyl halide carbon atom → lesser steric hindrance for the incoming nucleophile

Racemic mixture from S_N1 NS (+ why optically inactive)

- nucleophile can attact the trigonal planar carbocation carbon centre from either side of the plane with equal probability → both enantiomers are formed in equal amounts, forming racemic mixture (why racemic)
- optically inactive as each enantiomer rotates plane polarised light in opposite directions but same magnitude, ∴ their rotating powers cancel out



Inversion of configuration in S_N2 NS:

- nucleophile attacks from the backside of the alkyl halide C atom in S_N2 mechanism, which results in an inversion of configuration during the reaction

Relative reactivity of RCl, RBr, RI:

- Down the group, atomic radius of X increases → bond length of C−X increases
- ∴ bond strength decreases → ∵ lesser energy needed to break the weaker C−X bond
- rate of hydrolysis / reactivity increases (slowest) R−Cl < R−Br < R−I (fastest)

Unreactivity of aryl halides

1. the p orbital of Cl atom overlaps with the π orbital of benzene ring, causing the lone pair of e^- on Cl atom to delocalise into benzene ring → ∴ C-Cl bond has some double bond character
2. rear side of C-X bond is blocked by bulky benzene ring which provides steric hindrance
3. π electron cloud of benzene ring will repel lone pair of e^- of incoming nucleophile

main reason — 1 other reasons (if more mark) — 2, 3

Hydroxy Compounds

Undergo NS because:

- C-OH bond is polar because the C atom is bonded to a more electronegative O atom.
∴ the C atom with the −OH group is electron deficient & carries a partial +ve charge and is susceptible to nucleophilic attack
- Substitution as R-OH is saturated

Why phenol undergo ES very readily

- OH group in phenol is a strong EDG → activates the ring strongly by increasing the e^- density of the benzene ring → susceptible to electrophilic substitution.

Carbonyl compounds

Undergo nucleophilic addition because:

- O atom is more electronegative than C atom, making C=O bond polar and carbonyl C atom e^- deficient → C atom susceptible to nucleophilic attack
- addition as a small molecule can be added across the unsaturated C=O to form a single product

Carbonyl vs Alkenes reactivity

- both react similarly to saturate the double bond by undergoing addition
- carbonyl susceptible to nucleophilic attack as C atom e^- deficient ∵ bonded to electronegative O atom
- no electronegativity difference between 2 C atoms in C=C bond → C=C bond non-polar and no electron deficient C atom for nucleophilic attack + π bond makes C=C bond e^- rich → repels approach of nucleophile and susceptible to electrophilic attack

Ketone vs Aldehyde reactivity

↳ Ketone generally less reactive to nucleophilic attack than aldehydes

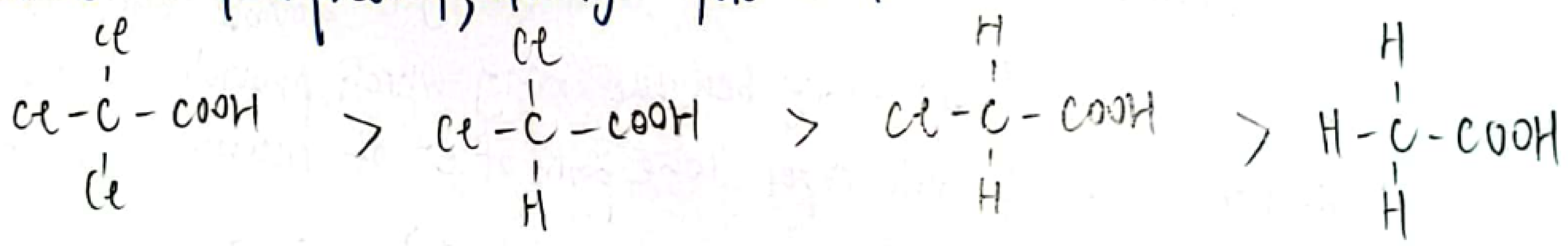
- presence of more e^- donating alkyl groups → carbonyl C atom in ketone less e^- deficient than that in aldehyde → less susceptible to nucleophilic attack
- presence of more bulky alkyl group → carbonyl C atom experiences more steric hindrance than aldehyde → prevent nucleophiles from attacking carbonyl C atom

Carboxylic acids

Chlorine-substituted ethanoic acids

1. No. of substituent groups

- more Cl atoms present, stronger the acid



$$\mathrm{Cl-\underset{Cl}{\overset{Cl}{C}}-COOH > Cl-\underset{H}{\overset{Cl}{C}}-COOH > Cl-\underset{H}{\overset{H}{C}}-COOH > H-\underset{H}{\overset{H}{C}}-COOH}$$

- Cl is e^- withdrawing group → more Cl atoms, the -ve charge on carboxylate ion is dispersed to greater extent → stabilise the conjugate base
- POE for acid dissociation shifts right → carboxylic acid more acidic → stronger acid

2. Position of substituent group

- closer the Cl atom to the carboxylate group, greater the e^- withdrawing effect → -ve charge on the carboxylate ion is more dispersed → stabilise conjugate base to a greater extent
- POE for acid dissociation shifts right → carboxylic acid more acidic → stronger acid

Why diprotic acids $K_{a_2} < K_{a_1}$

- Harder to remove a positively charged proton from a negatively charged ion due to electrostatic attraction

Nitrogen compounds

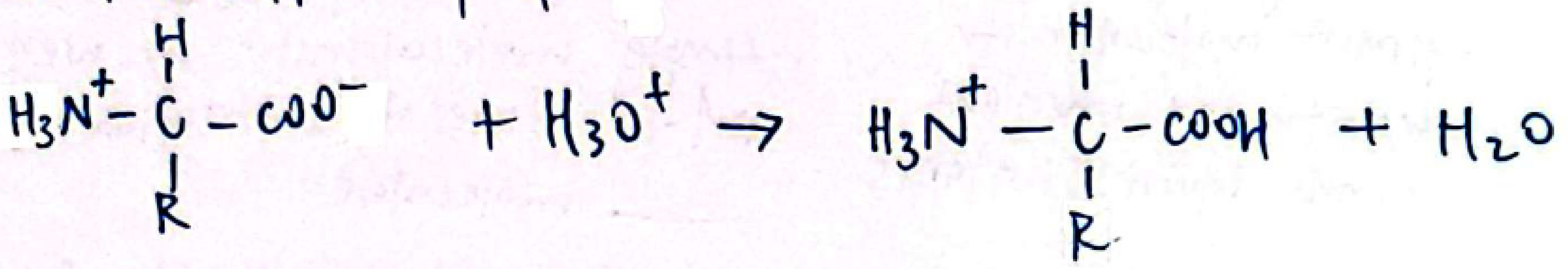
Basicity of 1°/2°/3° amines

- $R_3N > R_2NH > RNH_2 > NH_3$ for base strength
  - e⁻ donating alkyl group increases e⁻ density on N atom → ↑ availability of lone pair of e⁻ on N atom to form dative bonds with proton → base str. increases

How amino acids acts as buffer

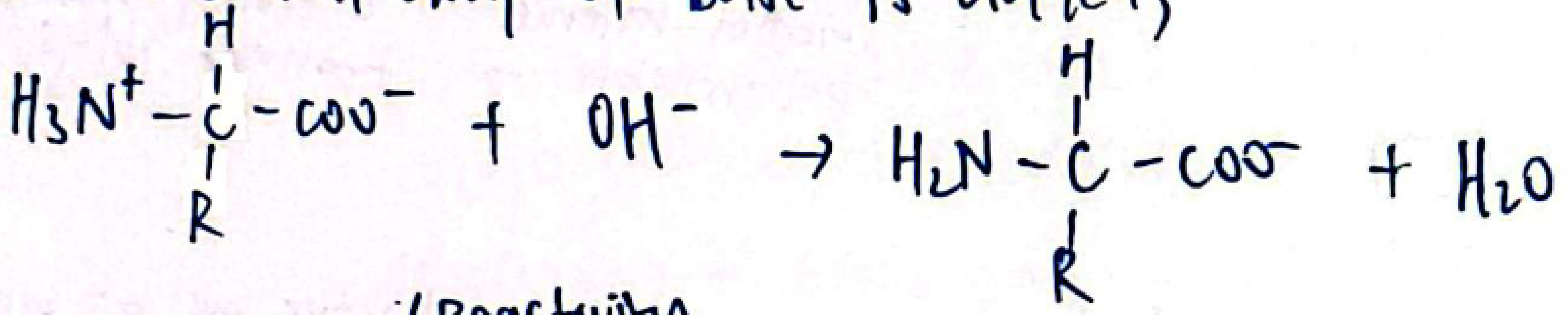
- aqueous solution of amino acids contains both $-COO^-$ and $-NH_3^+$ (zwitterion) groups, ∴ α-amino acids are amphoteric. since buffer is able to resist large changes in pH when small amt of acid / base is added to it,

- when small amt of acid is added,



$$H_3N^+-\overset{H}{\underset{R}{C}}-COO^- + H_3O^+ \rightarrow H_3N^+-\overset{H}{\underset{R}{C}}-COOH + H_2O$$

- when small amt of base is added,



$$H_3N^+-\overset{H}{\underset{R}{C}}-COO^- + OH^- \rightarrow H_2N-\overset{H}{\underset{R}{C}}-COO^- + H_2O$$

Acidity / Basicity / Reactivity comparison

case 1: orbital overlap

- p orbital of X/O/N atom overlaps with π orbital of benzene ring / C=O / C=C
- lone pair of e⁻ in X/O/N delocalises into benzene ring / C=O / C=C

case 2: substituent

1. e⁻ donating / withdrawing group increases / decreases the electron density on the C/N atom
2. acyl chloride bonded to 2 highly electro-ve Cl and O atoms while alkyl halide bonded to only one electro-ve Cl atom → acyl chloride C atom more e⁻ deficient → more susceptible to Nu attack

Acidity: -ve charge on O atom more effectively dispersed over ___ / dispersed into benzene rin
→ stabilises the conjugate base to greater extent → POE of dissociation of acid lies more to the right → donates H⁺ more readily → stronger acid

Basicity: ↑ availability of lone pair of e⁻ on N atom to form dative bond w/ a proton → base strength increases

Period 3

- Ionic radius : ↓ from Na^+ to Si^{4+} and ↓ from P^{3-} to Cl^-
  - nuclear charge increase and shielding effect remain approx. the same
  - effective nuclear charge ↑, net attraction between nucleus and outermost e^- ↑
  - r(cations) < r(atomic) as cations have 1 less electron shell
  - r(anions) > r(atomic) as anions have more e^- than protons, ∴ net attractive force on valence e^- lesser
  - r(anions) > r(cations) as anions has 1 additional electron shell

- Melting point

· high mp, ↑ from Na to Al · giant metallic str, strong metallic bonds btwn cations and sea of delocalised valence e^- · mp ↑ from Na to Al due to ↑ metallic bond str as ↑ no. of valence e^-	· Si highest mp · giant molecular str w/ strong covalent bonds btwn Si atoms	· Low mp, $S_8 > P_4 > Cl_2 > Ar$ · simple molecular str w/ weak id-id forces of attraction btwn ___ molecules · mp ↓ in order S_8, P_4, Cl_2, Ar as larger e^- cloud → e^- cloud more polarisable → stronger id-id

- Electrical conductivity

· high conductivity, ↑ from Na to Al as ↑ no. of delocalised valence e^- to act as charge carriers.	· low conductivity as Si is a metalloid/ semiconductor	· non-conductor as e^- are localised in covalent bonds ∴ no delocalised valence e^- to act as charge carrier

Chloride Reactions:

- with water

	$NaCl$ (s)	$MgCl_2$ (s)	$AlCl_3$ (s)	
Structure	giant ionic	giant ionic	simple molecular	
Reaction	hydration to form ions	hydration + slight hydrolysis (due to high charge density)	hydration + hydrolysis	* little amt of H_2O: $AlCl_3(s) + 3H_2O(l) \rightarrow Al(OH)_3(s) + 3HCl(g)$
	$NaCl(s) + aq \rightarrow Na^+(aq) + Cl^-(aq)$	$MgCl_2(s) + 6H_2O(l) \rightarrow [Mg(H_2O)_6]^{2+}(aq) + 2Cl^-(aq)$ $[Mg(H_2O)_6]^{2+}(aq) + H_2O(l) \rightleftharpoons [Mg(H_2O)_5(OH)]^+(aq) + H_3O^+(aq)$	$AlCl_3(s) + 6H_2O(l) \rightarrow [Al(H_2O)_6]^{3+} + 3Cl^-(aq)$ $[Al(H_2O)_6]^{3+}(aq) + H_2O(l) \rightleftharpoons [Al(H_2O)_5(OH)]^{2+}(aq) + H_3O^+(aq)$	
pH	7	6.5	2-3	

	$SiCl_4(l)$	$PCl_5(s)$
Structure	simple molecular	simple molecular
Reaction	hydrolysis	hydrolysis
	$SiCl_4(l) + 2H_2O(l) \rightarrow SiO_2(s) + 4HCl$	$PCl_5(s) + 4H_2O(l) \rightarrow H_3PO_4(aq) + 5HCl(aq)$ *cold water : $PCl_5(s) + H_2O(l) \rightarrow POCl_3(aq) + 2HCl(g)$
pH	1-2	1-2

Oxide reactions :

	$Na_2O(s)$	$MgO(s)$	$Al_2O_3(s)$	$SiO_2(s)$	$P_4O_{10}(s)$	$SO_3(g)$
Structure	Giant ionic	Giant ionic	Giant ionic	Giant molecular	Simple molecular	Simple molecular
Acid/Base nature	Basic	Basic	Amphoteric	Acidic	Acidic	Acidic
pH of solution	13	9 (sparingly soluble)	7 (insoluble)	7 (insoluble)	1-2	1-2

$\underbrace{\text{MgO, } Al_2O_3}_{\because \text{ strong ionic bonds}}$

↳ strong covalent bonds

Reactions w/ water :

- $Na_2O(s) + H_2O(l) \rightarrow 2NaOH(aq)$
- $MgO(s) + H_2O(l) \rightleftharpoons Mg(OH)_2(aq)$
- $P_4O_{10}(s) + 6H_2O(l) \rightarrow 4H_3PO_4(aq)$
- $SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$

Reactions w/ acid :

- $Na_2O(s) + 2HCl(aq) \rightarrow 2NaCl(aq) + H_2O(l)$
- $MgO(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l)$
- $Al_2O_3(s) + 6HCl(aq) \rightarrow 2AlCl_3(aq) + 3H_2O(l)$
- $Al(OH)_3(aq) + 3HCl(aq) \rightarrow AlCl_3(aq) + 3H_2O(l)$

Reactions w/ base :

- $Al_2O_3(s) + 2NaOH(aq) + 3H_2O(l) \rightarrow 2NaAl(OH)_4(aq)$
- $Al(OH)_3(aq) + OH^-(aq) \rightarrow [Al(OH)_4]^-(aq)$
- $SiO_2(s) + 2NaOH(conc) \rightarrow Na_2SiO_3(aq) + H_2O(l)$ → will only react w/ conc. NaOH AT HIGH TEMP.
- $P_4O_{10}(s) + 12NaOH(aq) \rightarrow 4Na_3PO_4(aq) + 6H_2O(l)$
- $SO_3(g) + 2NaOH(aq) \rightarrow Na_2SO_4 + H_2O(l)$

Why oxidation number of oxides ↑ across period:

- no of valence e^- available for bond formation ↑

Why P and S show several oxidation no. :

- they can expand their octet as they contain vacant and energetically available 3d orbitals

Group 2

- Reactivity / reducing strength :
  - ↑ from Mg to Ba → seen by the more negative $E^{\theta}(M^{2+}/M)$ values → M more readily oxidised
  - explanation: ionisation energy explanation
- Reaction with water
  - Be: No reaction w/ cold water or steam ∵ of protective layer of BeO (Be diagonal r/s w/ Al, properties similar to Al_2O_3)
  - Mg: slowly w/ cold water, vigorously w/ steam

    $Mg + H_2O \rightarrow MgO + H_2$

    $MgO(s) + H_2O(l) \rightleftharpoons Mg(OH)_2(aq)$
  - Ca to Ba: rapidly w/ cold water, w/ ↑ vigour down the group

    $M + 2H_2O \rightarrow M(OH)_2 + H_2$

Oxides :

	BeO	MgO	CaO SrO BaO
Nature	ionic and amphoteric	→	ionic and basic
Reaction w/ water	insoluble ∵ strong ionic bond	$MgO + H_2O \rightleftharpoons Mg(OH)_2$ slightly soluble ∵ strong ionic bond	↑ reactivity w/ water $MO + H_2O \rightarrow M(OH)_2$
pH	7	8–9	10 – 13
Reaction w/ acid / base	acid: $BeO + 2HCl \rightarrow BeCl_2 + H_2O$ base: $BeO + 2NaOH + 2H_2O \rightarrow Na_2Be(OH)_4$		all react w/ acid to give salt + water $MO + 2HCl \rightarrow MCl_2 + H_2O$

Thermal decomposition of nitrate / carbonate / hydroxide

- $M(NO_3)_2 \rightarrow MO + 2NO_2 + \frac{1}{2}O_2$
- $MCO_3 \rightarrow MO + CO_2$
- $M(OH)_2 \rightarrow MO + H_2O$

Thermal stability

- down the group, ionic radius ↑, charge density ↓
- cations lower polarising power → e^- cloud of NO_3^- / CO_3^{2-} less polarised
- ∴ N–O / C–O bond less weakened, more energy needed to break bond → thermal stability ↑ dtg.

Group 17

- Volatility
  - all volatile, volatility ↓ from Cl_2 to I_2.
  - ∵ strength of id-id btwn molecules ↑ → size of e^- cloud ↑ → e^- cloud more polarisable

- Reactivity / Oxidising strength
  - ↓ from Cl_2 to I_2 as seen by the less positive $E^\circ(X_2/X^-)$ values, X_2 less readily reduced
  - explanation : ↓ electronegativity (std expl^n) + ↓ how strongly nucleus attract incoming e^-

- Reaction w/ thiosulfate, $S_2O_3^{2-}$
  - Cl_2 and Br_2: $4X_2 + S_2O_3^{2-} + 5H_2O \rightarrow 8X^- + 2SO_4^{2-} + 10H^+$
  - I_2: $I_2 + S_2O_3^{2-} \rightarrow 2I^- + S_4O_6^{2-}$

- Thermal Stability
  - $2HX(g) \rightarrow H_2(g) + X_2(g)$
  - Down the group, atomic radius ↓ → lower degree of orbital overlap between H and X atoms
  - Bond strength and therefore bond energy decreases
  - Less energy req. to break H-X bond → thermal stability decreases.

  HCl low tendency to decompose
  HBr decomposes slightly
  HI decomposes readily

Transition Elements

- A transition element is a d-block element which forms one or more stable ions with a partially filled d subshell
- A ligand is a molecule or anion that has one or more lone pair of electrons for dative bond formation w/ a transition metal atom or ion
- A complex ion is a central metal atom/ion surrounded by molecules or anions which forms dative bonds to the metal centre
- Coordination number indicates the number of dative bonds around the central atom or ion in the complex

Atomic radius / 1st IE from s block to d block
- ↑ nuclear charge from Ca to Sc → e^- added to inner 3d subshell, !-shielding of 4s e^-
- shielding from the 3d e^- is ineffective → ENC↑ → net attraction btwn nucleus and outermost e^- ↑ → . . .

Atomic radius / 1st IE across period
- ↑ nuclear charge as no. of protons ↑ → additional e^- added to inner 3d orbitals → shielding effect ↑ → ENC remain approx. constant → net attraction btwn nucleus and outermost e^- approx constant → . . .

- 3^{rd} IE and higher
  - ↑ across period → 3d e^- are now outermost e^- → nuclear charge ↑ as proton no. ↑ → shielding effect remain approx. the same → ENC ↑ → net attraction btwn nucleus and outermost e^- ↑ → 3^{rd} IE ↑ across period.

- Melting point
  - higher than s-block elements, mp > 1000°C
  - both 3d and 4s e^- can be delocalised as close in energies → form stronger metallic bonds than s-block elements like Ca which has only 2 valence e^- for metallic bond formation

- Density
  - Denser than s-block, gradually ↑ across period
  - $\rho = \frac{mass}{volume}$ → smaller atomic radius + larger atomic mass

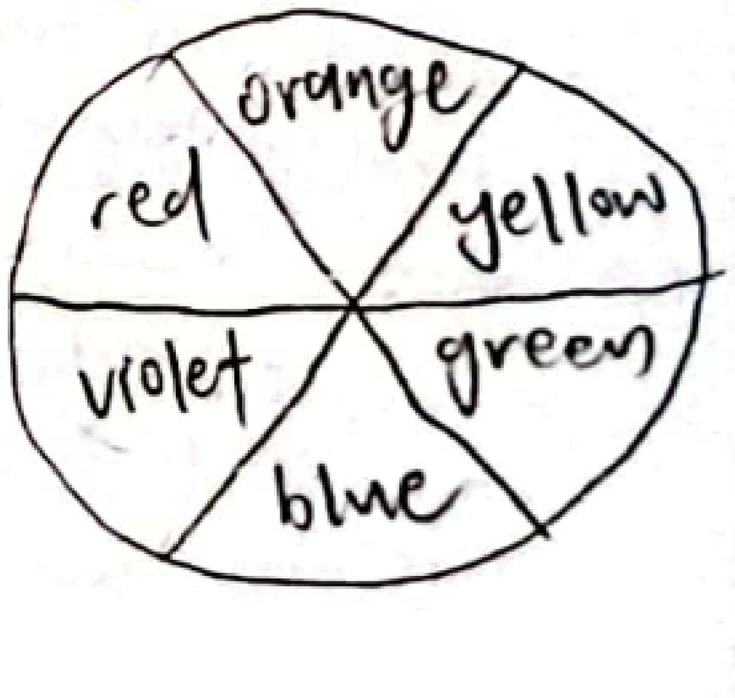
- Electrical conductivity
  - Higher than s-block, gradually ↑ across period
  - Both 3d and 4s e^- are delocalised as they are close in energies → more available charge carriers

- Variable oxidation states
  - Transition elements can lose different numbers of 4s and 3d e^- as they are close in energies ⇒ max os = no. of 4s + no. of unpaired 3d e^-. After Mn, pairing of ~~of~~ e^- occurs
  - s-block metal readily lose all 4s e^- to gain noble gas configuration → subsequent removal of e^- is from inner quantum shell → req. large amt of energy ∴ fixed o.s.

*CrO_4^{2-} ions formed instead of Cr^{6+}
- Formation of Cr^{6+} will require very high IE
- $Cr^{6+}(aq)$ will also have very high charge density and high polarising power → will polarise water ligands to form CrO_4^{2-}

- Formation of complex ions
  - Transition elements have vacant 3d orbitals of low energy levels that can accomodate lone pair of e^- donated by ligands (also ∵ metal ion have high charge density)
  - s-block metal do not have vacant 3d orbitals of compatible energy to accomodate the lone pairs (also ∵ it has low charge density)

- Colour of complex ions
  - In an octahedral complex ion, presence of ligands causes the energy level of the five 3d orbitals to split into two different energy levels
  - Energy difference, ΔE corresponds to wavelengths of light in visible spectrum
  - when light energy corresponding to ΔE is absorbed, an electron is promoted from d orbital of lower energy to one of higher energy
  - unabsorbed wavelengths are transmitted, colour observed is complementary to colour absorbed



Explanation for Im an
- has partially filled d-orbital so that d-d transitions are possible

- s-block metal is white/colourless as electronic transition of 4s e^- to 4p orbital involve absorption of energy outside the visible region

Crystal Field Theory / why there is splitting of energy levels

- $d_{x^2-y^2}$ and d_{z^2} orbitals have lobes directed along the x, y, z axes pointing towards approaching ligands while d_{xy}, d_{yz}, d_{xz} orbitals have lobes that are directed between the axes along which the ligands approach
- ∴ e^- in the $d_{x^2-y^2}$ and d_{z^2} orbitals experience stronger repulsions from the lone pair of e^- in ligands than those in d_{xy}, d_{yz} and d_{xz} orbitals
- $d_{x^2-y^2}$ and d_{z^2} are higher energy d-orbitals, while d_{xy}, d_{yz} and d_{xz} are lower energy d-orbitals

END of notes (for now? hopefully?)

Soluble complexes

- $[Ag(NH_3)_2]^+$
- $[Zn(OH)_4]^{2-}$
- $[Al(OH)_4]^-$
- $[Zn(NH_3)_4(H_2O)_2]^{2+}$
- $[Cu(NH_3)_4(H_2O)_2]^{2+}$
- $[CuCl_4]^{2-}$
- $[Cr(OH)_6]^{3-}$