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COLLEGEBOARD SAT II CHEMISTRY

GUIDE & SUMMARIZED NOTES ON THE SYLLABUS

TABLE OF CONTENTS

FORMAT & STRATEGIES

1. TYPES OF QUESTIONS

1.1 Part A

- A set of choices is provided and the next few questions refer to these choices.
- Each choice may be used one, more than once, or not at all in each set

1.2 Part B

- Every question contains two statements, I and II.
- For each question, decide whether I is true or false and whether II is true or false
- Finally, fill in oval CE only if statement II is a correct explanation of II
- It is possible that I and II are both true however II is not the explanation of I
- These type of questions will be in answered in a special section of the answer sheet labelled "chemistry" and the questions will begin from 101

1.3 Part C

- These are general five-choice questions with each set of options independent from the previous
- Questions may contain words in capital such as NOT, LEAST or EXCEPT and therefore are not always looking for the correct option
- Questions could also be to fill in a gap of a statement

2. PROBLEM SOLVING PROCESS

1) **Clarify the problem:** separate problem into facts, conditions, questions that need to be answered and establish the goal
2) **Explore:** examine sufficiency of data, organize data and apply knowledge, skills and understanding
3) **Select a strategy:** choose a method to solve the problem. Not each one can be applied to every problem and most of the times you will use a combination than one exclusively: Strategies: trial-and-error, reduction, working backwards, knowledge based
4) **Solve:** apply skills to carry out chosen strategy
5) **Review:** examine reasonableness of solution through estimation and evaluating effectiveness of process

STRUCTURE OF MATTER

1. ATOMIC STRUCTURE

- **Atom**: smallest particle of an element that retains the chemical properties of the element
- **Element**: substance that cannot be broken down into simpler substances through chemical reaction

1.1 Subatomic Particles

Subatomic Particle	Relative Charge	Relative mass/ a.m.u
Protons (P)	+1	1
Neutrons (n)	0	1
Electrons (e^-)	-1	1/1837

1.2 Protons, Neutrons and Electrons

- Nucleus of atom contains protons and neutrons
- Nucleus surrounded by electron shells
- **Isotopes**: atoms of the same element that have different number of neutrons in the nucleus
- Reading the symbol:

Nucleon number → 11
Proton number → 5
B

- **Isoelectronic ions:** ions having same no. of e^-s
- **Ground state:** lowest energy state for particle
- **Excited state:** any energy level above the ground state

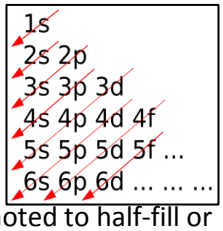
1.3 Electron Configuration

- Electrons are arranged in energy levels called shells
- Each shell is described by a principle quantum no. (P.Q)
- As the P.Q. increases, energy of shell increases
- **Orbital:** region in space where there is a maximum probability of finding an electron
- Each orbital can hold 2e^-s in opposite directions
- When e^-s are placed in a set of orbital of equal energy, they occupy them singly and then pairing takes place
- e^-s placed in opposite direction, creating a spin to reduce repulsion
- Completely filled or half-filled are more stable

1.4 Subshells

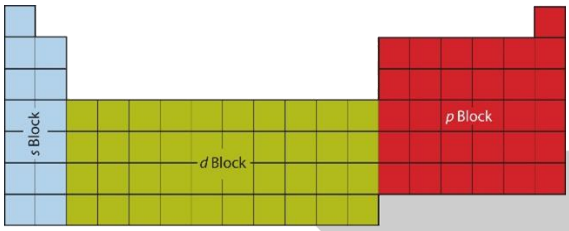
	s	*p*	*d*	*f*
Orbitals	1	3	5	7
Max e^-s	2	6	10	14

- **Aufbau's principle:** method of showing how atomic orbitals are filled in a definite order to give lowest energy arrangement possible



- Energy difference between $4s$ & $3d$ very small ∴ an e⁻ from $4s$ can be promoted to half-fill or full-fill $3d$ orbital, to make atom more stable
- When filling, fill 4s before 3d and when removing, also remove first from 4s

1.5 Periodic Table



Acid Properties Increase →

← Base Properties Increase

Atomic Radii Decrease →

Ionization Energy Increase →

1.6 Groups and Properties of Metals

- **General Properties:**
  - Malleable, ductile, lustrous
  - Oxidise (rust & tarnish) to form positive ions (cations)
  - Conduct heat an electricity
- **Alkali Metals:**
  - Most reactive metal family
  - Form alkaline solutions when reacted with water
  - Reactivity increases down group
- **Alkali Earth Metals:**
  - Also reactive but less in comparison to 1st group
  - Reactivity increases down the group
- **Transition Metals:**
  - Usually form coloured solutions
  - Have several possible oxidation states
  - Act as catalysts
  - Silvery-blue at r.t.p (except copper and gold)
  - Solid at r.t.p (except mercury)
  - Magnetic

1.7 Groups and Properties of Non-Metals

- **General Properties:**
  - Do not conduct electricity
  - All elemental gases are included here
  - Hydrogen placed with metals, is also a non-metal
- **Halogens:**
  - Exist as diatomic molecules
  - Most commonly used to form salts
  - Reactivity decreases down the group

1.8 Atomic Radius Trends

Down a Group	Across a Period
INCREASES	*DECREASES*
• New shells added • Attraction of nucleus to valence e⁻s decreases	• Shell no. remains same • Proton no. increases • Effective nuclear charge increases

- **Ionic radius:** describes size of ions
  - **Metals:** lose electrons ∴ lose a shell and ionic radius less than atomic radius
  - **Non-metals:** gain electrons ∴ gain a shell and ionic radius greater than atomic radius

1.9 General 1st I.E Trends

Down a Group	Across a Period
DECREASES	*INCREASES*
• Atomic radius increases • Effective nuclear charge decrease	• Atomic radius decreases • Shielding effect constant

1.10 Electronegativity

- **Electronegativity:** strength with which the atoms attract valence electrons in a chemical bond. Scale of 0 to 4, 4 being greatest electronegativity
- **Across a period:** increases
- **Down a group:** decreases

1.11 Radiations

	α-particle	β-particle	γ-ray
Identity	Helium nucleus	Fast-moving electron	Electro-magnetic
Symbol	$^{4}_{2}He$	$^{0}_{-1}e$	γ
Charge	$+2$	-1	0
Relative Mass	4	$\frac{1}{1840}$	0
Speed	Slow	Fast	V of Light
Energy	Discrete	Varying	
Stopped by	Paper	Few mm of aluminium	Few cm of lead
Range in air	5cm	12m	No Specific
Ionizing power	High	Low	Very Low

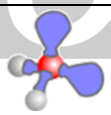
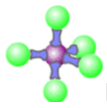
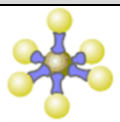
1.12 Nuclear Reactions

Types of Decay	Particle	Charge	Change in Mr	Change in Ar
Alpha Decay	α	2+	-4	-2
Beta Decay	β	1-	0	+1
Gamma Radiation	γ	0	0	0
Positron Emission	β^+	1+	0	-1
Electron Capture	e^-	1-	0	-1

- **Law of conservation of matter:** in all forms of decay neither protons nor neutrons are created or destroyed
- There are 2 main types of nuclear reactions:
  - **Fusion:** 2 smaller nuclei form a larger nucleus
  - **Fission:** a large nucleus splits into 2 smaller nuclei

2. Molecular Shapes

2.1 VSEPR and Hybrid Theory

Type	Shape	Angle	Hybrid	Example
2 Pairs of e⁻s				
2B	Linear	180°	sp	 CO_2
3 Pairs of e⁻s				
3B	Trigonal Planar	120°	sp^2	 BF_3
4 Pairs of e⁻s				
4B	Tetrahedral	109.5°	sp^3	 CH_4
3B 1L	Pyramidal	107°	sp^3	 NH_3
2B 2L	Angular	104.5°	sp^3	 H_2O
5 Pairs of e⁻s				
5B	Trigonal Bipyramid	90°	sp^3d	 PF_5
6 Pair of e⁻s				
6B	Octahedral	90°	sp^3d^2	 SF_6

B = Bonded Pair

L = Lone Pair

3. Chemical Bonding

- **Chemical Bonds:** result of the attraction between atoms, ions or other particles
- **Octet Rule:** all atoms try to achieve the closest noble gas electron state

3.1 Ionic Bonding

- Ionic bond is the electrostatic attraction between oppositely charged ions.
- Elements of the ions have a difference in electronegativity of greater than 1.67
- Structure: giant ionic lattice, crystalline solids
- Have high melting and boiling points

3.2 Covalent Bonding

- Covalent bond is the bond formed by the sharing of pairs of electrons between two atoms.
- Covalent compounds are made of molecules which are held together by weak intermolecular forces
- They have low melting and boiling points
- Sharing occurs unequally due to different electronegativities of the 2 atoms in the bond
  - *Nonpolar:* difference between 0 & 0.4
  - *Polar:* difference between 0.4 & 1.67

3.3 Metallic Bonding

- Strong electrostatic forces of attraction between metal cations and delocalized mobile electrons
- **Structure:** lattice of +ve ions surrounded by mobile e⁻s
- Mobile electrons can conduct electricity and heat

3.4 Properties and Structures

	Metal + Non-Metal	Metals Only	Non-Metals Only		
Bonding	Ionic	Metallic	Covalent		
			Molecular		Macromolecule
			Polar	Non-Polar	
Structure	Giant ionic lattice	Giant metallic lattice	Molecular Structure		Giant Covalent
Particles Present	+ve and -ve ions	+ve ions and -ve electrons	Molecular		Atoms
Forces	Electrostatic		Weak intermolecular and Strong intramolecular		Strong covalent
M.P. / B.P.	High		Low		Highest
Solubility in Water	Yes	No	No except hydrogen-bonded		No
Physical State at R.T.P	Solid (hard)		Soft solid, liquid or gas		Very Hard Solid
Electrical Conductivity	Molten/ aqueous	Good	No		No except graphite

3.5 Intermolecular Forces

- **Intermolecular forces:** weak forces present between two covalent molecules

London Dispersion Forces (Induced Dipole-Dipole):

- Very weak forces present between **non-polar molecules**
- Due to constant motion of e^-s, at an instant, a non-polar molecule develops poles due to distortion of electron density giving rise to instantaneous dipole, which is able to induce a dipole in the adjacent molecules

Permanent Dipole-Dipole Forces:

- Two or more neutral molecules attract due to their partial charges
- Unlike charges creates a strong attractive force
- Greater difference in polarity = stronger forces

Hydrogen Bonding:

- Strongest type of intermolecular force in covalent bonds
- For hydrogen bonding to occur, we need:
  - Molecule having a H atom bonded to F, O or N
  - Molecule having F, O or N atom with lone pair of e^-s

STATES OF MATTER

1. GASES

1.1 Kinetic-Molecular Theory

- Matter in all forms is composed of small particles
- Particles of matter are in constant motion; solids vibrate about a fixed position, liquids slide past each other and gases are in continuous and random motion
- Collisions of particles with themselves and the walls are elastic; no loss of energy

1.2 Ideal Gas Laws

- Gas molecules move rapidly and randomly
- Distance between gas molecules is greater than diameter of molecules ∴ volume is negligible
- No forces of attraction/repulsion between molecules
- All collisions between particles are elastic E_K conserved
- Temperature of gas related to average E_K of molecules
- **Conditions at which gases behave ideally:**
  - **High** temperature
  - **Low** pressure

Limitations of Ideal Gas Laws:

- Real gases do not obey kinetic theory in two ways:
  - There is **not** zero attraction between molecules
  - We **cannot** ignore volume of molecules themselves

Deviations visible at low temp. & high pressure

- Molecules are close to each other
- Volume of molecules not negligible relative to container
- VDW forces present, pulling molecules to each other
- Pressure is lower than expected from ideal gas
- Effective volume is less than expected from ideal gas

1.3 Measuring Pressure of Gases

Unit	Abbreviation	= to 1 atm
Atmosphere	atm	1 atm
Millimetres of Hg	mm Hg	760mm Hg
Torr	torr	760 torr
Pascals	Pa	101,000 Pa

- **Standard temp. and press. (STP):** 273K and 1 atm

1.4 Gas Laws

- **Graham's Law of Effusion (Diffusion):** rate of effusion of a gas is inversely proportional to the square root of its molecular mass

$$\frac{Rate\ A}{Rate\ B} = \frac{\sqrt{Molecular\ Mass\ of\ B}}{\sqrt{Molecular\ Mass\ of\ A}}$$

- **Charles's Law:** at constant pressure, the volume of a gas varies directly with the absolute temperature

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \text{ or } \frac{V}{T} = k$$

- **Boyle's Law:** at constant temperature, the volume of a gas varies inversely as pressure changes

$$P_1V_1 = P_2V_2 \text{ or } PV = k$$

- **Guy-Lussac's Law:** at constant volume, the pressure varies directly with the absolute temperature

$$\frac{P_1}{T_1} = \frac{P_2}{T_2} \text{ or } \frac{P}{T} = k$$

- **Combined Gas Law:**

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

- **Dalton's Law:** the pressure of a mixture of gases is equal to the sum of partial pressures of the component

$$P_{total} = P_1 + P_2 + \cdots + P_x$$

- **Ideal Gas Law:**

$$PV = nRT$$

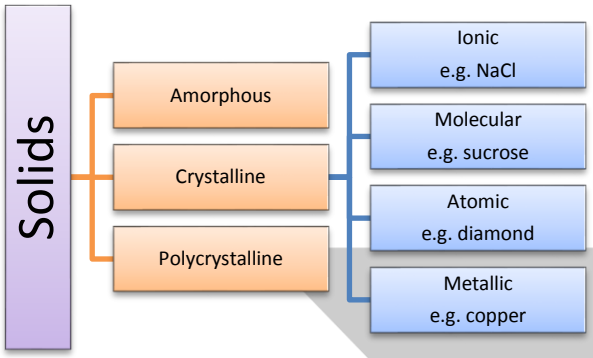
2. LIQUIDS AND SOLIDS

2.1 Solids

- Particles held together by ionic or string covalent bonds:
  - Attractive force is very strong
  - Definite shape and volume

Types of Solids:

- **Amorphous:** solids without form, random structure with little or no long-range ordering e.g. glass
- **Crystalline:** regular structure, particles in a repeating pattern
- **Polycrystalline:** solid with a large number of small crystals in which structure is regular but crystals arranged in randomly



2.2 Liquids

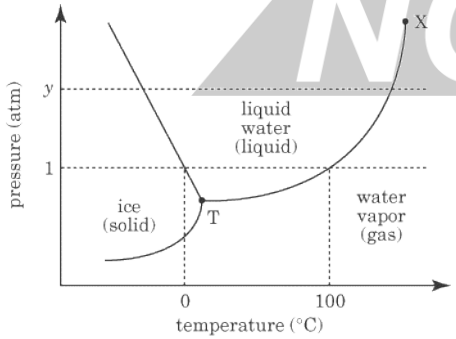
- Particles have strong forces of attraction
  - Particles have more freedom than in solid
  - Indefinite shape but definite volume

2.3 Phase Changes

- **Heat of fusion:** energy put in to melt a substance
- **Heat of vaporisation:** energy needed to cause transition from liquid to gas

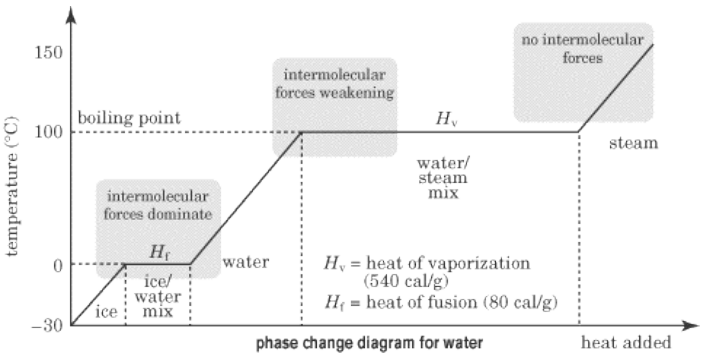
2.4 Phase Change Diagram

1st type: temperature against pressure



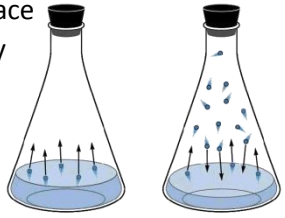
- **Triple point:** temperature and pressure at which the substance can exist in all three phases in equilibrium (T)

2nd type: energy against temperature (heating curve)



2.5 Phase Equilibrium

- **The evaporation of a liquid in a closed container**
  - Constant evaporation from surface
  - Particles continue to break away from surface but are trapped in space above the liquid
  - As gaseous particles collide, some of them hit the surface of the liquid again, and become trapped there
  - An equilibrium is set up in which number of particles leaving surface is balanced by number re-joining it.



liquid water molecules ⇌ vapour water molecules

  - In this equilibrium, there will be a fixed number of the gaseous particles in the space above liquid.
- **Vapour pressure:** pressure exerted by a vapour in equilibrium with a liquid.
- Vapour pressure **increases** as:

Temp. ↑ → Ek ↑ → Forces weaken → More vapor

3. Solutions

3.1 Measuring Concentration

- **Molarity (M):** no. of moles of solute per litre of solution

$$M = \frac{moles\ of\ solute}{litres\ of\ solution}$$

- **Molality (m):** no. of moles of solute per kg of solvent

$$m = \frac{moles\ of\ solute}{kg\ of\ solvent}$$

- **Percent by Mass:** another expression of concentration

$$Percent\ by\ Mass = \frac{mass\ of\ solute}{mass\ of\ solute\ \&\ solvent} \times 100$$

3.2 Solubility

- **Rule of solutions:**
  - Polar solvent dissolves polar solute
  - Non-polar solvent dissolves non-polar solute
- **Solubility:** degree to which a given solute will dissolve in a given solvent
- **Solubility of solids in water:**
  - Increases with increasing temperature
  - Not affected by pressure
- **Solubility of gases in water:**
  - Decreases with increasing temperature
  - Increases with increasing pressure

3.3 Electrolyte

- Ionic substances dissolve, ionic bonds broken and substance dissociates into mobile ions
- **Electrolyte:** solutions capable of carrying conducting electrical current

3.4 Colligative Properties

- Properties of solutions that depends on the ratio of solute to solvent particles: freezing and boiling point, and vapour pressure
- **Freezing Point Depression (FPD):**
  - Pure solvent has a certain freezing point and when freezing particles must cluster.
  - In solution, solute particles get in the way and prevents tight clustering
  - Causes the freezing point to decrease
- **Boiling Point Elevation (BPE):**
  - Pure solvent has a certain boiling point and when boiling the particles try to escape.
  - In solution, solute particles get in the way which requires more energy to escape
  - Causes the boiling point to increase
- **FPD and BPE expression:**

$$\Delta T = kmi$$

where k is a constant dependent on solvent, m is molarity and i is the no. of ions in solution (1 for covalent compounds)

- FPD and BPE depend only on type of solvent and number of solute particles, not type of solute particles

Reaction Types

- **Synthesis:** 2 or more reactants form 1 product

$$A + B \rightarrow C$$

- **Decomposition:** breakdown into 2 or more products

$$AB \rightarrow A + B$$

- **Displacement:** units replace each other
  - **Single:** more reactive element replaces less reactive

$$A + BC \rightarrow AC + B$$

  - **Double:** 2 compounds react to form 2 new compounds

$$AB + CD \rightarrow AD + CB$$

- **Combustion:** hydrocarbon burnt in oxygen

$$C_xH_y + O_2 \rightarrow CO_2 + H_2O$$

- **Hydrolysis:** reaction with water

$$X^- + H_2O \rightarrow XH + OH^-$$

1. Acids and Bases

1.1 Acid and Bases Theories

	Acid	Base
Arrhenius	Produces H_3O^+	Produces OH^-
Brønsted-Lowry	Proton donor	Proton acceptor
Lewis	Accepts electrons	Donates electrons

1.2 Strong and Weak Acids and Bases

- **Strong acids/bases:** acids/bases which dissociate almost completely in solutions

$$HX_{(aq)} \rightarrow H^+_{(aq)} + X^-_{(aq)}$$

$$HX_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + X^-_{(aq)}$$

- **Weak acids/bases:** acids/bases which are only partially dissociated in solutions

$$X^-_{(aq)} + H^+_{(aq)} \rightarrow XA_{(aq)}$$

$$X^-_{(aq)} + H_2O_{(l)} \rightarrow HX_{(aq)} + OH^-_{(aq)}$$

- Strong and weak acids and bases can be distinguished by the pH value of their aqueous solutions
- **Amphoteric:** substance that can act as an acid or a base
- **Monoprotic acids:** donate one H^+ proton per molecule
- **Diprotic acids:** donate two H^+ protons per molecule

1.3 Properties of Acids

- Acids are conductors of electricity in aqueous solution
- Acids will react with metals that are more reactive than hydrogen ions to liberate hydrogen gas
- Acids react with bases to neutralize each other, forming water and a salt
- Acids react with carbonates to release carbon dioxide

1.4 Strong Acids

- HCl hydrochloric acid
- HBr hydrobromic acid
- HI hydroiodic acid
- HNO_3 nitric acid
- H_2SO_4 sulfuric acid
- $HClO_4$ perchloric acid

1.5 Properties of Bases

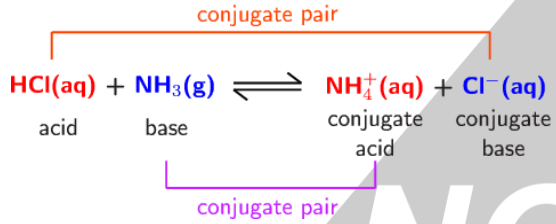
- Bases are conductors of electricity in aqueous solution
- Bases react with acids to neutralize each other, forming water and a salt
- Bases react with fats to form soaps

1.6 Strong Bases

- KOH potassium hydroxide
- NaOH sodium hydroxide
- $Ba(OH)_2$ barium hydroxide
- $Sr(OH)_2$ strontium hydroxide
- $Ca(OH)_2$ calcium hydroxide

1.7 Conjugate Acids and Bases

- When acid-base reacts, an equilibrium mixture is formed



- The stronger an acid, the weaker its conjugate base
- The stronger a base, the weaker its conjugate acid

1.8 Acid/Base Properties of Salts

- Weak acid + strong base = basic salt
- Weak acid + weak base = neutral salt
- Strong acid + weak base = acidic salt
- Strong acid + Strong base = neutral salt

1.9 Acid/Base Equilibrium

- When an acid/base is added to water, it dissociates and a dynamic equilibrium is set up between acid/base and ions
- Calculating acid dissociation constants:

$$K_a = \frac{[H^+][X^-]}{[HX]}$$

- Calculating base dissociation constants:

$$K_b = \frac{[X^+][OH^-]}{[XOH]}$$

1.10 pH Scale

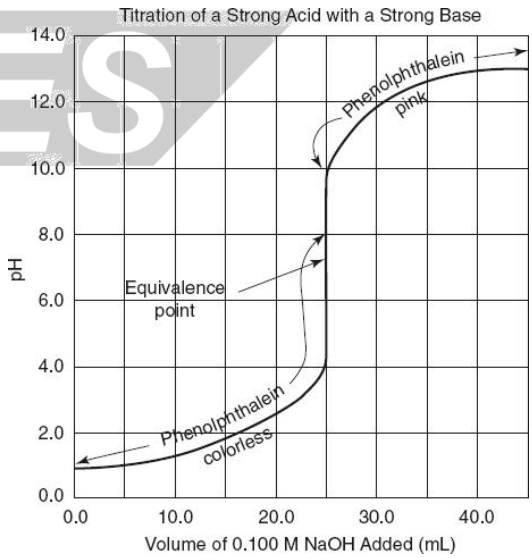
- Gives strength of acid/bases by concentration of H^+/OH^- ions

$$pH = -\log[K_a] \qquad pOH = -\log[K_b]$$
$$pH = 14 - pOH$$

- **Calculating the pH:**
  - Firstly, use all the information and calculate the constant as you would in other equilibrium questions
  - Find the log without a calculator, see the power of 10 it is to and that is your log
  - If it is K_a, this is the pH
  - If it is K_b, do (14 – Ans) hence finding pH

1.11 Titration

- Adding strong acid or base of known identity and concentration (**titrant**) to unknown acid or base solution with an indicator
- Equivalence (end) point; enough titrant has been added to neutralize the subject acid or base
- **Titration curve:**



  - Equivalence point where curve steepest

1.12 Indicators

Indicator	pH Range	Color Below Lower pH	Color Above Higher pH
Methyl orange	3.1 – 4.4	Red	Yellow
Bromthymol blue	6.0 – 7.6	Yellow	Blue
Litmus	4.5 – 8.3	Red	Blue
Phenolphthalein	8.3 – 10.0	Colourless	Pink

- Titration of strong acid and strong base; end point will be pH 7 and any indicator can be used
- Titration of strong acid and weak base; use methyl orange; changes colour in acid region
- Titration of strong base and weak acid; use phenolphthalein; changes colour in basic region

1.13 Buffer Solutions

- Equilibrium systems that resist changes in acidity and maintain constant pH when acid or bases are added
- It is a mixture of a weak acid and its conjugate base or a weak base and its conjugate acid
- Work by reacting with any added acid or base to control the pH

2. Oxidation-Reduction

2.1 Redox Reactions

- **Oxidation:** loss of electrons, gain of +ve charge
- **Oxidising agent (OA):** reduces itself to oxidise others
- **Reduction:** gain of electrons, gain of –ve charge
- **Reducing agent (RA):** oxidises itself to reduce others
- **Half-reaction:** equation that shows only either oxidation or reduction
- **Redox reaction:** reaction where both oxidation and reduction occur
- Can be shown with changes in oxidation numbers of elements from the product side to the reactant side

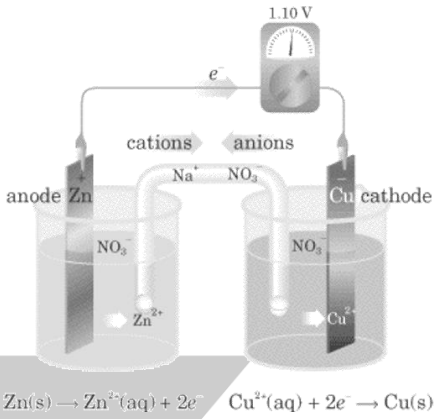
2.2 Calculating Oxidation Numbers

Ionic Molecules: group number = valence electrons

- **Rules:**
  - Atoms in a diatomic molecule; oxidation number = 0
  - Oxygen in a compound; oxidation number = -2
  - Oxygen as peroxide; oxidation number = -1
  - 1st group elements & hydrogen; oxidation number = +1
  - H with highly reactive metal; oxidation number = -1
- Following these rules, all other atoms in a covalent bond must balance out the charge

2.3 Voltaic (Galvanic) Cells

- Separates the reduction and oxidation reaction and forces the electrons to flow over a wire (producing electricity) from the oxidation reaction (at the anode) to the reduction reaction (at the cathode).



$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$

- **Anode:** +ve electrode, attracts –vely charged anions
- **Cathode:** -ve electrode, attracts +vely charged cations
- **Salt bridge:** device that maintains electrical neutrality
- The EMF of a voltaic cell is due to potential energy difference of the electrons before and after the transfer
- **Standard Reduction Potentials (E^o):** potential that would be produced between a given half-reaction and hydrogen

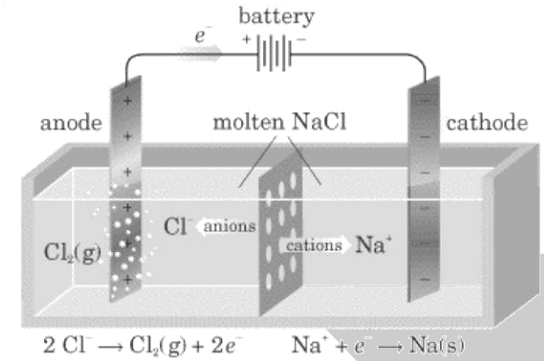
  EMF = cathode potential – anode potential
- +ve EMF indicates a spontaneous process
- The greater the difference between the electrode potentials, the larger the EMF

2.4 Standard Reduction Potentials

Element			Standard Half-Cell Voltages Reduction Electron Reaction	Electrode Potential, E^0
Potassium	↑	▬	$K^+ + e^- \rightarrow K$	–2.93 V
Calcium			$Ca^{2+} + 2e^- \rightarrow Ca$	–2.87 V
Sodium	Increasing	Increasing	$Na^+ + e^- \rightarrow Na$	–2.71 V
Magnesium	Tendency	Tendency	$Mg^{2+} + 2e^- \rightarrow Mg$	–2.37 V
Aluminum	for	for	$Al^{3+} + 3e^- \rightarrow Al$	–1.67 V
Zinc	Atoms	Ions	$Zn^{2+} + 2e^- \rightarrow Zn$	–0.76 V
Iron	to	to Gain	$Fe^{2+} + 2e^- \rightarrow Fe$	–0.44 V
Tin	Lose	Electrons	$Sn^{2+} + 2e^- \rightarrow Sn$	–0.14 V
Lead	Electrons	and	$Pb^{2+} + 2e^- \rightarrow Pb$	–0.13 V
Hydrogen	and	Form	$2H^+ + 2e^- \rightarrow H_2$	0.00 V
Copper	Form	Atoms	$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34 V
Mercury	Positive	of the	$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	+0.79 V
Silver	Ions	Metal	$Ag^+ + e^- \rightarrow Ag$	+0.80 V
Mercury			$Hg^{2+} + 2e^- \rightarrow Hg$	+0.85 V
Gold	▬	↓	$Au^{3+} + 3e^- \rightarrow Au$	+1.50 V

2.5 Electrolytic Cell

- Energy from a cell used to bring about a nonspontaneous redox reaction
- Used for:
  - Purification of elements
  - Separating metals
  - Charge batteries
- **Example:**



3. Precipitation

3.1 Precipitation Reaction

- Occur when cations and anions in aqueous solution combine to form an insoluble ionic solid; precipitate
- It is a double replacement reaction when the reactants are soluble and one of the products formed is insoluble.

3.2 Solubility Table

Soluble Salts	Insoluble Salts
All sodium, potassium and ammonium salts	The rest
All nitrates	N/A
Chlorides	Except silver and lead
Sulphates	Except barium, lead and calcium
Potassium, sodium and ammonium carbonates	All other carbonates

Stoichiometry

1. Mole Concept

1.1 The Mole

- **Mole:** amount of substance that has the same number of particles (atoms, ions, molecules or electrons) as there are atoms in exactly 12g of the carbon-12 isotope.
- **Avogadro's constant:** number of atoms, ions, molecules or electrons in a mole = 6.02×10^{23}
- **Molar mass:** mass of a given substance divided by its amount of substance; g/mol

1.2 Formulae

- **Empirical formula:** gives simplest ratio of different atoms present in a molecule
- **Molecular formula:** gives actual numbers of each type of atom in a molecule
- Molecular formula can be calculated using the **Mr** of a compound and its empirical formula

$$Molecular\ Formula = (Empirical\ Formula)_n$$

$$\text{Where } n = \frac{Molecular\ Mass}{Mass\ of\ Empirical\ Formula}$$

2. Chemical Equations

2.1 Balancing Equations

- Use correct formulas for all reactants and products
- Begin balancing with most complicated-looking group
- Polyatomic ions can be a single unit if unchanged
- Balance single element reactants & products last
- Keep your eye out for diatomic molecules
- If stuck, double most complex group & try again
- Make sure all coefficients are in lowest-possible ratio
- Numbers too complex, restart

2.2 Stoichiometric Calculations

$$Moles = \frac{Mass}{Molar\ Mass}$$

$$Volume\ of\ a\ Gas = Moles \times 24$$

- Formula applies to gases at r.t.p.
- Unit of volume is dm^3 and $1000cm^3 = 1dm^3$

$$Concentration = \frac{Moles}{Volume}$$

2.3 Reagents

- Reactants may not always be present in perfect number of reacting moles
- **Limiting reagent:** reactant with fewer number of reacting moles
- **Excess reagent:** reactant with greater number of reacting moles

2.4 Yields

- **Theoretical:** equations say we can get this much of product from that much of reactant
- **Actual:** experiments say we got this much of product from that much of reactant
- **Percent:** ratio of actual to theoretical as a percentage

$$\frac{Actual\ Yield}{Theoretical\ Yield} \times 100 = Percent\ Yield$$

EQUILIBRIUM & REACTION RATES

1. EQUILIBRIUM SYSTEMS

- **Reversible reaction:** a reaction in which products can be changed back to reactants by reversing the conditions
- **Dynamic Equilibrium:** the state of a reversible reaction carried out in a closed container where the rates of forward and backward reactions are equal and constant

1.1 Le Chatelier's Principle

- When a chemical system in dynamic equilibrium is disturbed (conditions changed) it tends to respond in such a way so as to oppose the change and a new equilibrium is set up

By Le Chatelier's principle, when the	temperature	is increased, the system opposes the change by favouring	the endothermic side
	pressure		the side with fewer gas mols
	reactant conc.		the forward reaction

1.2 Equilibrium Constants

Equilibrium Constant

Expressed in terms of **concentration**	Expressed in terms of **partial pressure**
$K_C = \frac{[Product]^{mols}}{[Reactant]^{mols}}$	$K_P = \frac{p(Product)^{mols}}{p(Reactant)^{mols}}$
Only liquids and gases	Only gases

- Large value of K_C/K_P ⇒ equi. towards products side
- Smaller value of K_C/K_P ⇒ equi. towards reactants side
- K_C/K_P changes only with changes in temperature
- Amount of reactants that disappear always appear in the products in same ratio as present in equation

2. RATES OF REACTIONS

- **Rate of a reaction:** change in concentration of reactants or products per unit time
- **Activation energy:** minimum energy colliding particles must possess for a successful collision to take place

2.1 Factors Affecting Rates of Reactions

Concentration: mole per volume of a substance

- Higher reactant conc. - forward reaction faster
- Higher product conc. – backward reaction faster
- Higher conc. = more chance of successful collisions

Temperature: thermal energy provided to molecules

- Higher temp. more E_K, increasing rate of reaction
- Lower temp. less E_K, decreasing rate of reaction

Catalyst: decreases E_a of reaction

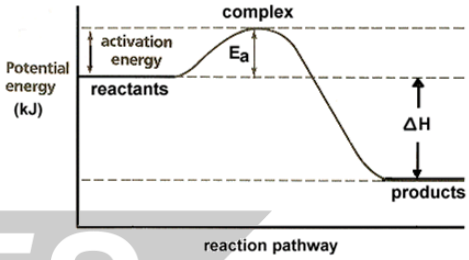
- Provides a different path for the reaction to take
- Unconsumed by reaction
- Biological catalysts are called enzymes

Physical State:

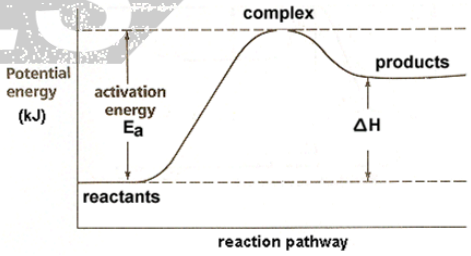
- *Homogenous:* reactants in same physical state
- *Heterogeneous:* reactants in different physical states
- In hetero; reactants limited to contact area therefore larger contact area allows for faster reaction

2.2 Potential Energy Diagrams

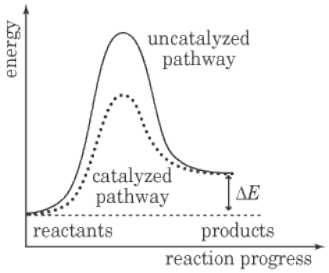
- **Exothermic reaction:**



- **Endothermic reaction:**



- **Addition of a catalyst:**



THERMOCHEMISTRY

1. CALORIMETRY & SPECIFIC HEATS

1.1 Heat Capacity

- Amount of heat energy required to raise temperature of a substance by 1°C
  - **Molar:** 1 mol of the substance
  - **Specific:** 1 gram of the substance

$$Q = mc\Delta T$$

1.2 Calorimetry

- Enthalpy of reaction measured using a calorimeter
- Measures heat flowing out/into system as reaction proceeds
- **Calorie:** amount of heat required to raise the temperature of 1 gram of a substance by 1°C

$$1\ calorie = 4.184\ joules$$

2. ENTHALPY

- **Enthalpy (ΔH):** heat content of a chemical reaction

2.1 Types

- **Enthalpy of Combustion (ΔH_{comb}):** 1 mole of element or compound is completely combusted under standard conditions in its standard state
- **Enthalpy of Formation (ΔH_f):** 1 mole of compound is formed from its elements under standard conditions in their standard states
- **Enthalpy of Fusion (ΔH_{fus}):** 1 mole of a solid melted to 1 mole of the corresponding liquid at the normal melting point of the substance
- **Enthalpy of Vaporisation (ΔH_{vap}):** 1 mole of a liquid changed to 1 mole of the corresponding gas at the normal boiling point of the substance

2.2 Hess's Law

- The total enthalpy change in a chemical reaction is independent of the route by which the chemical reaction takes place as long as the initial and final conditions are the same.
- **Reason to use Hess's Law:**
  - Std. conditions hard to maintain (e.g. exo/endo)
  - Elements don't always react directly

3. ENTROPY

- **Entropy (ΔS):** measure of the degree of disorder
- Universe favours a more disordered system; nature tends to chaos

3.1 Laws of Thermodynamics

- **1st Law:** energy can neither be created nor destroyed; only transformed from one type to another
- **2nd Law:** disorder of universe, its entropy, is constantly increasing
- **3rd Law:** entropy of a perfect crystal at 0K is zero

3.2 Rules when Determining Entropy

- Greater disorder in a system - larger entropy
- Entropy of substance always increases changing state from solid to liquid to gas
- When pure solid or liquid dissolves in solvent, entropy of substance increases
- When gas molecules escapes from solvent, there is an increase in entropy
- Entropy generally increases with increasing molecular complexity
- Reactions that increase no. of moles of particles often increase entropy of system

4. GIBBS FREE ENERGY EQUATION

$$\Delta G = \Delta H - T\Delta S$$

ΔH: enthalpy ΔS: entropy

T: temperature in Kelvin

- The sign of ΔG used to predict spontaneity of a reaction at constant temperature and pressure

4.1 Interpreting Sign of ΔG

- **ΔG negative:** reaction (probably) spontaneous
- **ΔG positive:** reaction improbable
- **$\Delta G = 0$:** system at equilibrium; no net reaction

4.2 Factors Affecting ΔG

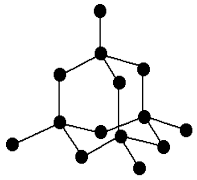
ΔH	ΔS	Will it happen?
-ve	+ve	Always
+ve	+ve	At high temp.
-ve	-ve	At low temp.
+ve	+ve	Never

DESCRIPTIVE & ORGANIC CHEMISTRY

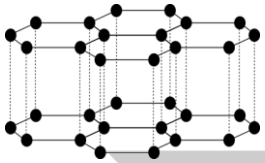
1. CARBON

1.1 Allotropes of Carbon

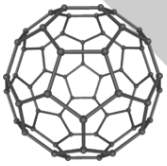
- **Carbon can exist as:**
  - **Diamond:** sp^3 hybrid orbitals



  - **Graphite:** sp^2 hybrid orbitals

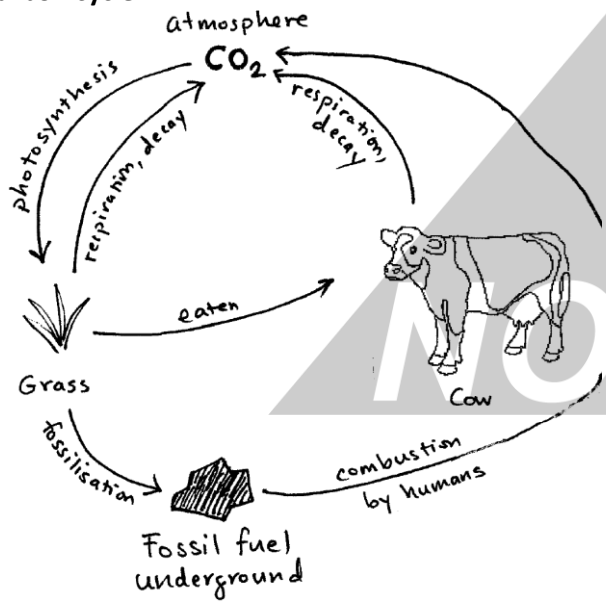


  - **Fullerenes:** spherical cage of carbons



1.2 Carbon Dioxide

- **Carbon cycle:**



- **Laboratory preparation:** react marble chips with acid

$$CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$$

- **Testing for carbon dioxide:** limewater turns milky

$$Ca(OH)_2 + CO_2 \rightarrow CaCO_3 + H_2O$$

Excess CO_2: continued passing will eliminate cloudy solution, forming soluble calcium hydrogen carbonate

$$CaCO_3 + H_2O + CO_2 \rightarrow Ca(HCO_3)_2$$

1.3 Uses of Carbon Dioxide

- Used to make carbonated drinks:

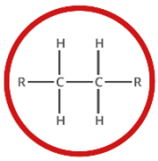
$$CO_2 + H_2O \rightarrow H_2CO_3$$

- Solid CO_2, "dry ice", used as refrigerant as it sublimes
- Used in fire extinguishers as it is heavier than air and prevents combustion from occurring
- Used by plants for photosynthesis:

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

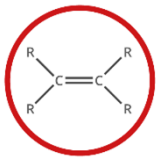
2. HYDROCARBONS

2.1 Alkanes



- Saturated; 4 single sp^3 bonds
- General formula = C_nH_{2n+2}
- Cycloalkanes: single bonded ring compounds; C_nH_{2n}
- First four alkanes = gases
- Non-polar; weak intermolecular forces of attraction
- Compounds separated due to difference in b.p. by fractional distillation

2.2 Alkenes



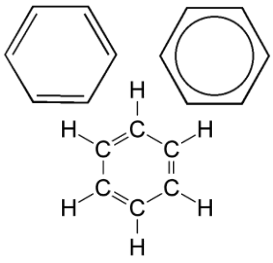
- Unsaturated; 3 sigma sp^2 bonds and 1 pi pure p bond on double bond
- General formula = C_nH_{2n}

2.3 Alkyne

$$R{-}C{\equiv}C{-}R'$$

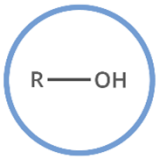
- Unsaturated; 2 sigma sp bond and 2 pi pure p bond on triple bond
- Name: -yne
- General formula = C_nH_{2n-2}

2.4 Aromatic



- Aromatic = unsaturated ring structures
- Benzene is the simplest aromatic compound = C_6H_6
- General formula = C_nH_{2n-6}

3. ALCOHOLS

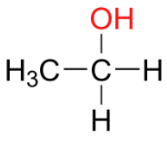
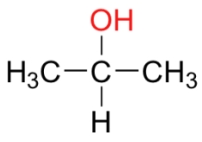
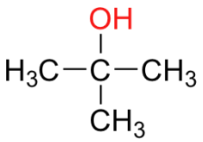


- Functional group: R–OH (hydroxyl)
- Colourless, flammable liquids

3.1 Uses of Alcohol

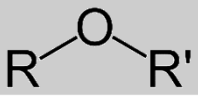
Methanol	Ethanol
• b.p. 65°C • Miscible with water • Poisonous; cause blindness if ingested • Used as fuel and solvent	• b.p. 78°C • Miscible with water and solvent for many substances • Used as fuel

3.2 Types of Alcohols

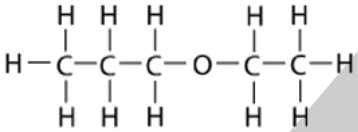
Primary 1°	Secondary 2°	Tertiary 3°
		

4. Ethers

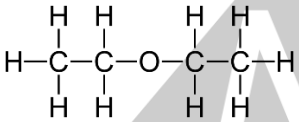
- Functional group: R–O–R'
- Formed when a primary alcohol is dehydrated with sulfuric acid



- **Naming ethers:**
  - Shorter chain becomes first part of name with –ane suffix changed to –oxy.
  - Longer alkane chain becomes suffix of name
  - E.g. ethoxypropane or ethyl propyl ether

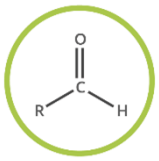


  - E.g. ethoxyethane or ethyl ether or diethyl ether



5. Carbonyl Compounds

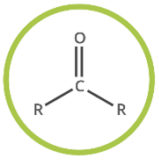
5.1 Aldehyde



- **Functional group:** R–CHO
- **Preparation from 1° alcohol:** mild oxidation of alcohol using oxidizing agent or inserting hot copper wire

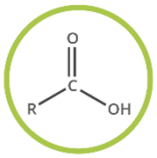
$$CH_3OH + [O] \rightarrow [CH_2(OH)_2] \rightarrow HCHO + H_2O$$

5.2 Ketone



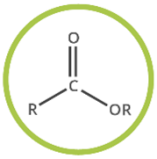
- Functional group R–CO–R'
- **Preparation from 2° alcohol:** oxidation of alcohol

6. Carboxylic Acids and Derivatives

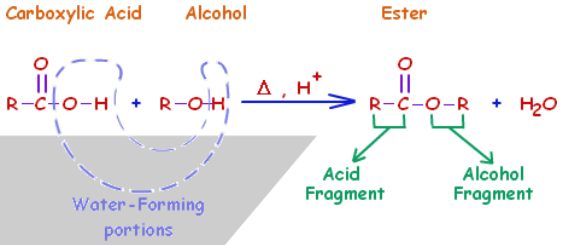


- Functional group: R–COOH
- **Preparation from an aldehyde:** mild oxidation of an aldehyde
- **Amino acids:** acids containing amine group (NH_2)

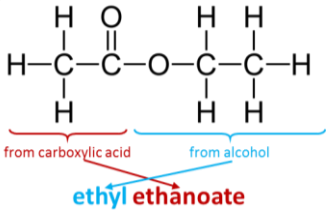
6.1 Esters



- Inorganic salts produced by reaction of alcohol and acid
- Reaction occurs with H_2SO_4

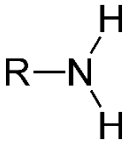


- Functional group: R–COO–R'
- Sweet smelling; used in perfumes and flavour extracts
- **Naming esters:**

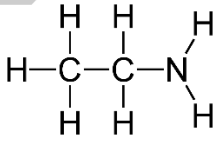


7. Amines and Amides

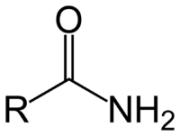
7.1 Amines



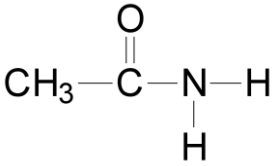
- **Functional group:** R–NH_2
- **Example:** ethanamine or ethylamine



7.2 Amides



- **Functional group:** RCO–NH_2
- Replaces hydrogen in the carboxyl group
- **Example:** ethanamide or ethylamide



8. Coloured Compounds

8.1 Soluble and Insoluble

Coloured Soluble Compounds	Coloured Insoluble Compounds
• Soluble copper salts are blue/green • Fe salts are red/brown • Cobalt salts are blue • Complex ions are often coloured	• AgCl is white • Chromate precipitates are orange • Dichromate precipitates are yellow • Hydroxide precipitates are white

8.2 Oxides

Oxide	Colour
Titanium oxide	White
Copper oxide	Green (patina)
Iron oxide	Red (rust)
Silver oxide	Black (tarnish)

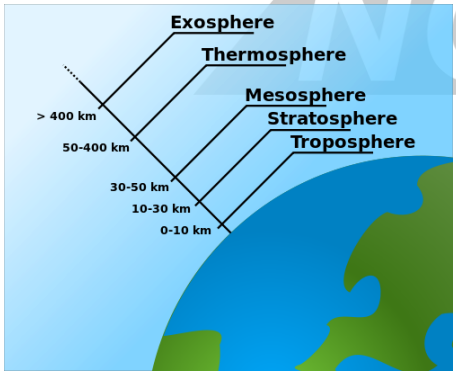
9. Environmental Chemistry

9.1 Earth's Atmosphere

Atmospheric Composition:

Nitrogen (N_2)	78%
Oxygen (O_2)	20%
Argon (Ar)	< 1%
Water Vapour	Variable
Other	< 1%

Layers of Atmosphere:



9.2 CFCs Effect on Ozone Layer

- CFCs escape in atmosphere and because of their inertness, remain without further reaction until they reach the stratosphere and ozone layer.
- In stratosphere, high energy U.V causes Cl atom to split of CFC molecule forming Cl· which reacts with ozone

9.3 Greenhouse Effect

- Refers to build up of CO_2 and other carbon gases in the atmosphere
- As Earth's surface absorbs solar radiation, it warms and radiates infrared radiation back into atmosphere
- Extra CO_2 reflects and traps radiation causing Earth to warm

9.4 Acid Rain

- Nitrogen and sulphur oxides produced by pollution interacts with water producing acid rain

$$SO_3 + H_2O \rightarrow H_2SO_4$$

$2NO_2 + H_2O \rightarrow HNO_3 + HNO_2$ or $NO_2 + H_2O + \frac{1}{2}O_2 \rightarrow HNO_3$

- Damages trees & plants, kills fish and other river life, buildings, statues and metal structures

The Laboratory

1. Laboratory

1.1 Technological Tools

- **Gravimetric balance:** direct readings to thousandth of a gram instead of mechanical balance
- **pH meters:** direct pH readings instead of indicators
- **Spectrophotometer:** measures % of light transmitted at specific frequencies so that molarity of sample can be determined without titration
- **Computer-assisted lab:** probes take readings and computer programs can monitor variables, collect data and produce graphs

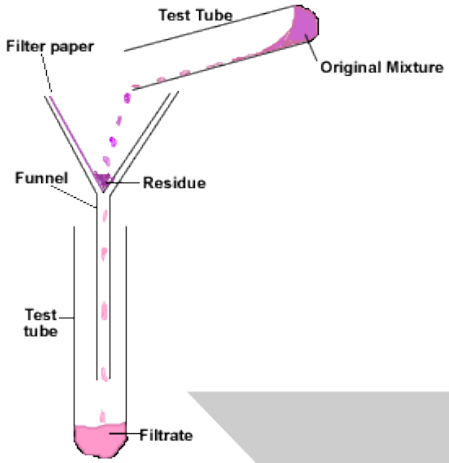
1.2 Laboratory Safety Rules

- Dress appropriately for lab; safety goggles, lab coat, tie back long hair and do not wear open-toed shoes.
- Know safety equipment available and how to use; eyewash fountain, fire blanket, fire extinguisher etc.
- Know dangers of chemical; read labels, don't taste/sniff
- Dispose of chemicals according to instructions.
- Always add acids and bases to water slowly to avoid splattering; can generate heat, steam, and splash
- Direct test tubes away when heating/reaction occurring
- Never pipette by mouth; don't use it as a suction pump
- Use the fume hood when dealing with toxic fumes
- Do not eat or drink in the lab; can ingest toxic chemicals
- Follow all directions; never haphazardly mix chemicals

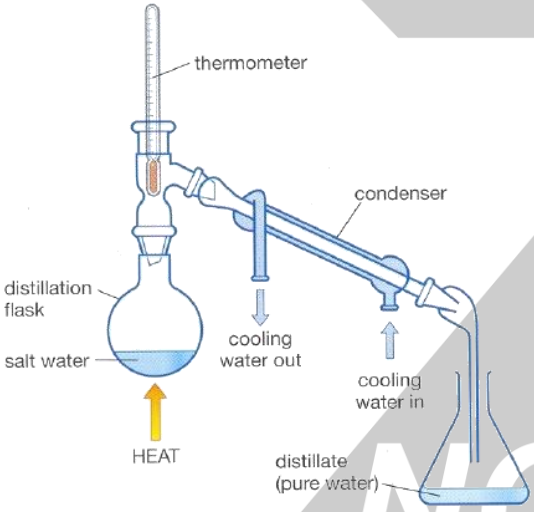
2. Experiment Setups

2.1 Methods of Separation

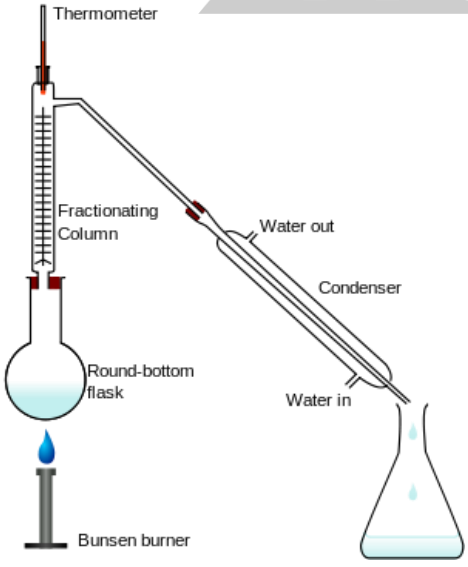
Filtration: separate solid from a liquid



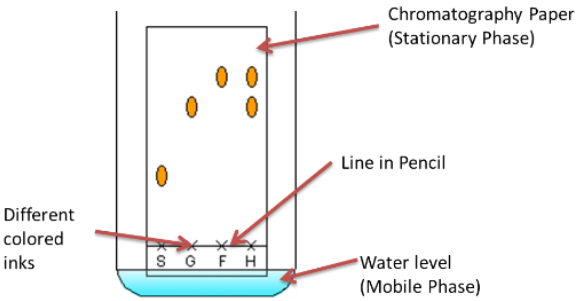
Simple distillation: separate solvent from a solution



Fractional distillation: separate liquids from each other

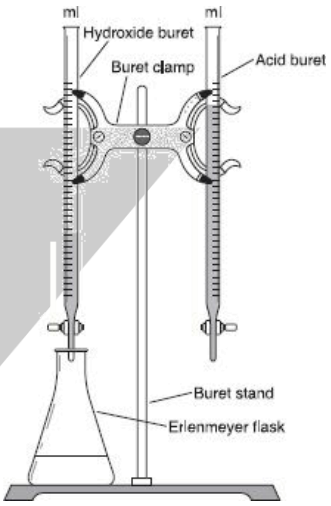


Chromatography: different substances from a solution

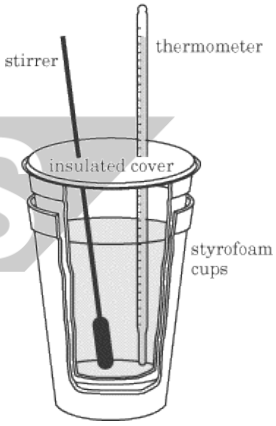


2.2 Other Set-ups

Titration: finding concentration of unknown acid or alkali



Calorimetry: finding heat capacity of a substance



3. Identifying Chemicals

3.1 Gas Tests

Gas	Test and result
Ammonia (NH_3)	1) Damp red litmus paper turns blue 2) White fumes form with HCl 3) Smell = sharp odour
Carbon dioxide	1) Bubble thru limewater; goes milky
Hydrogen (H_2)	1) Place light splinted; ignites, pop 2) Burns with blue flame

Gas	Test and result
Hydrogen Chloride (HCl)	1) Damp blue litmus turns red 2) Smell = choking odor
Hydrogen Sulfide (H_2S)	1) Damp lead acetate paper turns brown-black 2) Smell = rotten egg odor
Oxygen (O_2)	1) Insert glowing splint = relights 2) Nitric oxide gas turns red-brown

3.2 Cations Test

Ion	Test and result
Ammonium (NH_4^+)	Add strong base (NaOH) - produces ammonium
Iron(II) (Fe^{2+})	NaOH or Ammonia = green ppt.
Iron(III) (Fe^{3+})	NaOH or Ammonia = red-brown ppt.
Hydrogen (H^+)	Blue litmus turns red

3.3 Anions Test

Ion	Test and result
Acetate ($C_2H_3O_2^-$)	Add conc. H_2SO_4 and warm = vinegar odor released
Carbonate (CO_3^{2-})	Add HCl; limewater turns milky
Chloride (Cl^-)	Add nitric acid, then aqueous silver nitrate = white ppt. soluble in $NH_{3(aq)}$
Hydroxide (OH^-)	Red litmus turns blue
Sulfate (SO_4^{2-})	Add HCl then $BaCl_2$ = white ppt
Sulfide (S^{2-})	Add HCl then test gas with lead acetate paper = turns black-brown (H_2S)

3.2 Colours of Solutions

Ion Present	Solution color
Cu^{2+}	Blue
Fe^{3+}	Yellow to orange (rusty)
Ni^{2+}	Green
MnO_4^-	Purple
CrO_4^{2-}	Yellow
$Cr_2O_7^{2-}$	Orange

3.3 Flame Tests

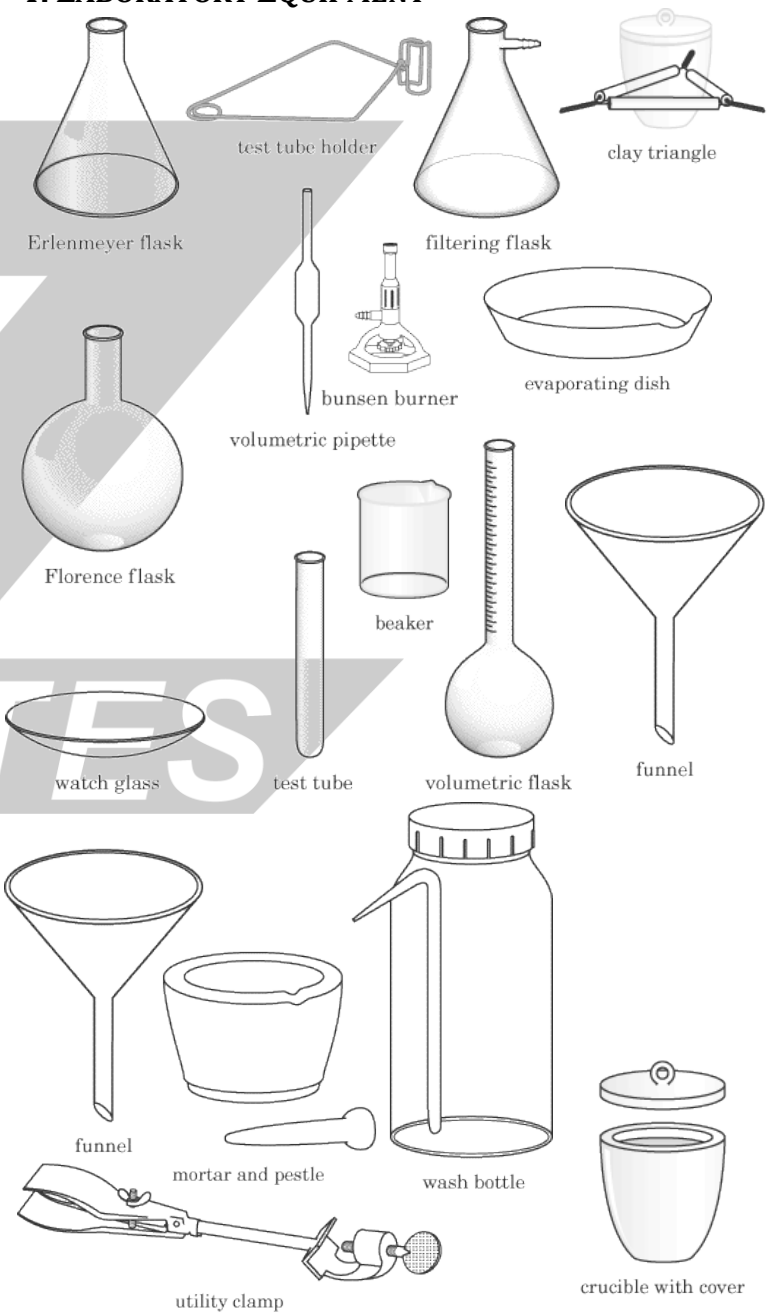
Metal	Flame Colour
Lithium, Strontium, Calcium	Red
Sodium	Yellow
Potassium	Lilac
Barium	Green
Copper	Blue-Green
Iron	Gold

3.4 Hydrogen Sulphide Tests

Bubble H_2S gas through solution being tested; ppt. formed

Metal	Colour of Sulphide ppt.
Lead	Brown-black
Copper, Silver, Mercury, Nickel, Iron	Black
Cadmium, Arsenic	Yellow
Antimony	Orange
Zinc	White
Bismuth	Brown

4. Laboratory Equipment



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