

Secondary 3 EOY Chemistry Notes

Elements, Compounds, Mixtures and Particulate Nature of Matter

Nature and Classification of Matter

Air

Metals

Kinetic Theory

The States of Matter

Phase Transitions

Heating Curve

Experimental Design and Purification Techniques

Measuring Gases

Measuring Liquids

Purification Techniques

Purity Indicators

Identification of Gases

Atomic Structure

Atoms

Isotopes

Ionization Energy

SPDF Electronic Configuration

Chemical Bonding and Structure

Ionic Bonds

Covalent Bonds

Dipoles

Intermolecular Forces of Attraction (IMF)

Metallic Bonding

Bonding, Structures and Properties

Chemical Formulae and Equations

Types of Chemical Formula

Common Ions

Soluble vs Insoluble Substances

Naming Ions and their Acids

Stoichiometry and the Mole

Acid Base Chemistry

Acid Base Reactions

Strength, Concentration and Basicity

pH of a Solution

Types of Oxides

Acid-Base Titration Curves

Common Acids and Alkalis

Observations of Reactions

Rules of Solubility in ppt reactions

Ionic Equations

Redox

Oxidation States

Redox Reactions

Oxidising agents and reducing agents

Chemical Energetics / Enthalpy

Energy conversion in reactions

Exothermic and Endothermic reactions

Reaction Kinetics

Factors affecting rate of reaction

Methods of increasing rate of reaction

Catalysts

Reaction Rate Graphs

Secondary 4 FE Chemistry Notes

[Periodic Table](#)

[Structure of the Periodic Table](#)

[Metals vs Non-metals](#)

[Trends in the Periodic Table](#)

[Trends within Groups](#)

[Group I vs Transition Metals](#)

[Qualitative Analysis](#)

[General Techniques in QA](#)

[Types of Reactions](#)

[QA of Anions](#)

[QA of Cations](#)

[QA of Gases](#)

[Chemical Equilibrium](#)

[Reversible Reactions](#)

[Dynamic Equilibrium](#)

[Le Chatelier's Principle](#)

[Metal Reactivity](#)

[Basic Reactivity](#)

[Reactions of Metals and Thermal](#)

[Stability in the Reactivity Series](#)

[Corrosion of Metals](#)

[Metal Alloys](#)

[Metal Recycling](#)

[Electric Cells](#)

[Fuel Cell](#)

[Electrolysis](#)

[Electrolysis](#)

[Electrolytic Cells](#)

[Factors affecting Electrolysis](#)

[Electrolytic Purification](#)

[Aluminum Purification](#)

[Copper Purification](#)

[Electroplating](#)

[Organic Chemistry](#)

Elements, Compounds, Mixtures and Particulate Nature of Matter

Nature and Classification of Matter

- Elements: Substance made of one type of atom which cannot be chemically split into simpler substances, metal or nonmetal.
- Compounds: Substance that contains two or more elements chemically combined, ionic or covalent
- Mixtures: Substance that contains two or more substances physically together but not reacted with each other chemically, made of elements and/or compounds

	Elements	Compounds
Definitions	A substance that is made up of only one type of atom, It cannot be chemically split into simpler substances.	A substance that contains two or more elements chemically joined together.
Similarities	<ul style="list-style-type: none"> • Both are made up of atoms, which are the building blocks. • Both are pure substances. • Both have fixed melting and boiling points. 	
Differences	<ul style="list-style-type: none"> • Made up of only one type of atom. • Can be metallic or non-metallic. • Cannot be chemically split into simpler substances. 	<ul style="list-style-type: none"> • Made up of more than one type of atom. • Can be formed by combining a metallic element with a non-metallic element, or several non-metallic elements. • Can be decomposed into its elements or into simpler compounds.

	Compounds	Mixtures
Similarities	<ul style="list-style-type: none"> • Both are made up of more than one type of atom. 	
Differences	<ul style="list-style-type: none"> • Can be separated into its elements by chemical methods. • Has a fixed melting and boiling point. • A chemical change takes place when a compound is formed. • The elements are always combined in a fixed proportion. • The properties of a compound differ from those 	<ul style="list-style-type: none"> • Can be separated into its components by physical methods. • Does not have a fixed melting or boiling point. • No chemical changes take place when a mixture is formed. • The proportion of the constituents in a mixture can vary. The elements are always combined in a fixed proportion by mass in a compound.

	of its constituent elements.	<ul style="list-style-type: none"> The properties of a mixture are the same as those of its constituents.
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Air

- 78% Nitrogen
- 21% Oxygen
- 0.9% Argon
- Note: Carbon Dioxide is not one of the most common. It is only 0.03%.

Metals

- Solid (except mercury) at rtp
- Shiny, good conductors of electricity and heat.
- Ductile (they can be drawn into thin wires).
- Malleable (they can be easily hammered into very thin sheets).

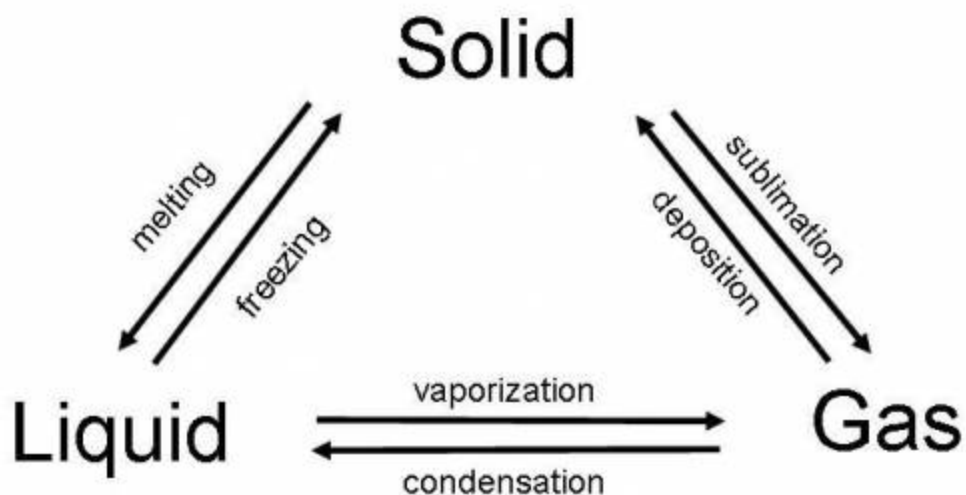
Kinetic Theory

- All matter made of atoms
- All atoms carry energy of motion (heat) in vibrational and translational movements
- Gas pressure is due to particle collisions
- Volume of particles in gases are negligible

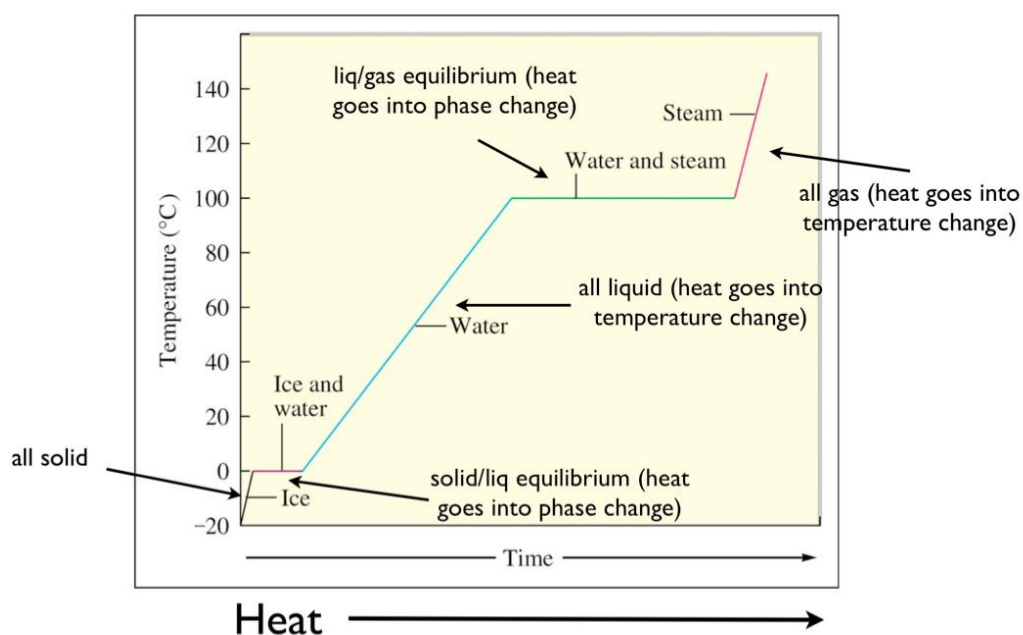
The States of Matter

	Solid	Liquid	Gas
Arrangement of Particles	Close together in an orderly arrangement	Close together but no orderly arrangement	Far apart and no orderly arrangement
Motion	Vibration	Vibration & Translation	Vibration & Translation
Shape and Volume	Fixed shape and volume	No fixed shape but has fixed volume	No fixed shape or volume
Compressible	No	No	Yes

Phase Transitions



Heating Curve



Experimental Design and Purification Techniques

Measuring Gases

- Typical Syringe
- Displacement of water if insoluble
- Upward/Downward delivery if less dense/denser than air
- Passed through drying agent if needed

Measuring Liquids

- Measuring Cylinder - 0dp
- Pipette - 1dp w/ fixed volume
- Burette - 1.2dp (2nd dp is 0 or 5)

Purification Techniques

- Dissolving, Filtering and Evaporating for insoluble and soluble
- Crystallization for dissolved salt
- Sublimation for high heat capacity + sublimable compound
- Separating Funnel for immiscible liquids
- Simple Distillation for single solvent
- Fractional Distillation for multiple solvents
- Reverse Osmosis for water and contaminants
- When describing process, have to use:
 - o Action (e.g. heat)
 - o Object (e.g. salt solution)
 - o Apparatus (e.g. evaporating dish)
 - o Reason / Theory (e.g. water evaporates, but salt doesn't)
 - o Result (e.g. salt is obtained)

Purity Indicators

- Sharp BP and MP
- Impurities BP+
- Impurities MP--

Identification of Gases

- Cl₂
 - o Moist blue litmus
 - o Green/Yellow gas evolved
 - o Litmus red > white (Cl₂ is bleaching agent)
- H₂
 - o Lighted splint
 - o Extinguishes /w 'pop' sound
- O₂
 - o Glowing Splint
 - o Splint Relights
- SO₂
 - o Filter paper + acidified KMn(VII)O₄ OR KCr(VI)₂O₇
 - o Choking gas
 - o Paper purple > colourless
- CO₂
 - o Bubble through lime water (Ca(OH)₂)
 - o White ppt of CaCO₃
- NH₃

- o Moist red litmus
- o Pungent gas
- o Litmus Red > Blue

Atomic Structure

Atoms

Protons – Charge: +1, Relative mass: 1

Electrons – Charge: -1, Relative mass: 1/1836 or 0

Neutrons – Charge: 0, relative mass: 1

The atomic structure can be written as a symbol called the Nuclide Notation:

- Mass number (nucleon number): Number of protons + Number of neutrons
- Atomic number (proton number): Number of protons/ electrons (in an atom with no charge)

Isotopes

- Definition: Isotopes are atoms of the same element with the same number of protons but different number of neutrons
- Same chemical properties but different physical properties (e.g. masses and boiling points)
- Naturally occurring or produced by nuclear reactions

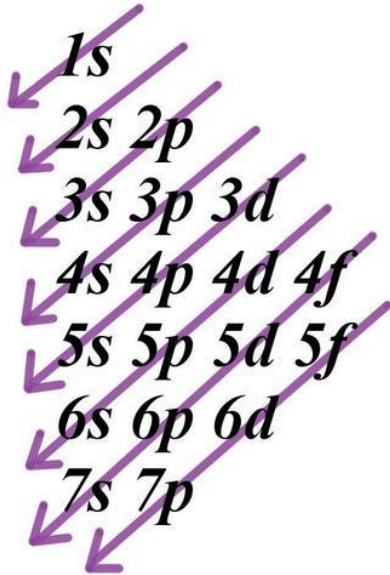
Ionization Energy

- Joules needed to “pluck off” on electrons
- Increases with lesser electrons remaining
- Large “jump” when crossing orbital/quantum shells

SPDF Electronic Configuration

Type of subshell	Number of orbitals
s	1
p	3
d	5
f	7

Order of filling orbitals:



Example:

Calcium $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

Note:

In xs^y , x is the shell, and y is the number of electrons

Each orbital contains 2 electrons

Chemical Bonding and Structure

Types of Bonds

- All chemical bonds are electrostatic in nature.
- Strong
 - Ionic
 - Covalent
 - Metallic
- Weak
 - Intermolecular Forces of Attraction (IMF)
 - Instantaneous dipole-induced dipole interaction (id-id)
 - Permanent dipole-permanent dipole interaction (pd-pd)
 - Hydrogen bonding

Ionic Bonds

- Electrons are “given and received”
- Occurs between metal and nonmetal atoms
- Metal atoms lose electrons to form cations
- Nonmetal atoms gain electrons to form anions

- Opposite charges of atoms attract each other
- Typically found in a giant crystal lattice structure

Covalent Bonds

- Electrons are “shared”
- Must involve nonmetal atoms, may or may not involve metal atoms
- Electrons contributed by atoms to form shared pairs, one pair creating one bond
- Can form single bonds, double bonds and triple bonds
- Macromolecules are molecules where a large number of atoms are joined by covalent bonds
- Polyatomic ions are typically made of covalent bonds

Coordinate Bonds

- Covalent bond in which the shared pair of electrons is provided by only one of the bonded atoms.
- Same properties as covalent bonds
- Coordinate bond is sometimes called a dative covalent bond.
- Conditions for coordinate bond to occur:
 - Donor atom must have at least 1 pair of unshared valence electrons
 - Acceptor donor must have at least one vacant orbital in its outer shell
- Examples: ammonium, carbon monoxide (work it out yourself)

Polar Covalent Bonds

- Electronegativity is a measure of the tendency of an atom to attract a bonding pair of electrons.
- When 2 atoms form a covalent bond, the more electronegative atom will pull the shared electrons toward itself, creating a partial negative charge. This leaves the other atom with a partial positive charge, creating a polar covalent bond
- If 2 atoms of equal electronegativity bond together, the covalent bond is nonpolar
- If 1 atom is a lot more electronegative than the other, the other atom will lose control of the shared electrons, and ions will be formed.

$0 \leq EN < 0.5$	Covalent
$0.5 \leq EN < 2.0$	Polar Covalent
$2.0 \leq EN < 4.0$	Ionic

Dipoles

- Separation of opposite charges.
- Thus, in a polar bond that connects one slightly positively charged atom to another slightly negatively charged atom, a dipole is present.

Intermolecular Forces of Attraction (IMF)

Instantaneous dipole-induced dipole (id-id) Interaction

- Electrons are in continuous, random motion, sometimes causing electrons to “clump” to one side of a molecule and creating an instantaneous dipole
- Instantaneous dipole will affect surrounding molecules, creating induced dipole
- Short lasting and weakest IMF
- Greater no. of electrons in molecules \Rightarrow greater id-id interaction
- Greater surface area of molecules \Rightarrow greater id-id interaction

Permanent dipole-permanent dipole (pd-pd) Interaction

- Polar covalent molecules and ionic structures have a permanent dipole
- Permanent dipole attracts opposite poles of other molecules
- If pd-pd is present, id-id is present

Hydrogen Bonding

- For hydrogen bonding to occur between 2 molecules, one molecule must have a hydrogen bonded to a N, O or F, and the other molecule must have N, O or F.
- How to draw: $X - H \cdots :Y$, where X and Y are N, O or F
- Hydrogen bonds may occur between the same type of molecules (in hydrogen fluoride) or between different types of molecules (e.g. between HF and H_2O) in aqueous hydrogen fluoride)
- Hydrogen bonds occur between molecules containing hydrogen covalently bonded to N, O or F. Such molecules have pd-pd and id-id interactions too. (Acronym: Full Of Nonsense)
- Hydrogen bonding $>$ id-id and pd-pd

Metallic Bonding

- Valence electrons all delocalised to form a “sea of electrons”
- Metallic atoms are held together by positive charged ions and negative charged sea of electrons
- Delocalised electrons are able to conduct heat and electricity well
- Metals are ductile and malleable as ions can “slide” across one another

Bonding, Structures and Properties

Type of structure	Simple Covalent	Macromolecular	Giant Ionic	Giant Metallic
Bonds Broken when heated	IMF	Covalent	Ionic	Metallic
Strong or Weak	Weak	Strong	Strong	Strong
Energy Required	Low	High	High	High
BP and MP	Low	High	High	High, except Group I metals

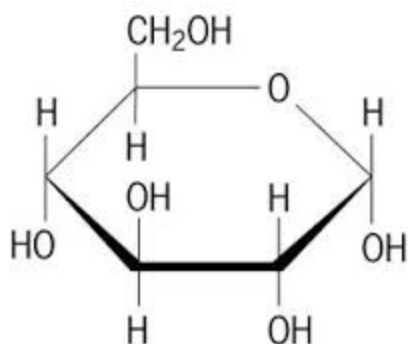
Conductivity	No except some polar covalents	No except Graphite	Molten or Aqueous	Yes
Solubility	No, unless in organic non-polar solvent	No	Yes	No, though reactive

[Substance] is a [structure type] which has [bond type], which are [property e.g. strength/polarity] and hence [property].

Chemical Formulae and Equations

Types of Chemical Formula

- Molecular formula is all atoms in a molecule: $C_6H_{12}O_6$
- Empirical formula is simplest ratio of atoms in a molecule: CH_2O
- Structural formula shows which atoms are bonded to what:



Common Ions

+	2+	3+	3-	2-	-
(S)Na (P)K (L)Li (A)NH ₄ (S)Ag (H)H	Most other metal ions	Al Fe(III)	N PO₄	CO₃ Cr ₂ O ₇ O SO₄ S SiO ₂	Cl ClO₃ I (M)MnO ₄ (O)OH (N)NO₃ NO ₂

Bolded Ions are -ate Ions

Soluble vs Insoluble Substances

Soluble substances (usually aq)	Insoluble substances (s)
All nitrate salts	-
Most Chloride salts	Lead and Silver
Most Sulfate salts	Lead, Barium and Calcium
Sodium, Potassium and Ammonium	Most Carbonate salts
Sodium, Potassium and Ammonium	Hydroxides and Oxides
Most common acids and alkalis	...NIL, unless otherwise stated in exam paper

Naming Ions and their Acids

Number of Oxygen Atoms	Ion	Acid
1 more than -ic	Per-ate	Per-ic
-ic	-ate	-ic
1 fewer than -ic	-ite	-ous
2 fewer than -ic	Hypo-ite	hypo-ous
No Oxygen	-ide	hydro-ic

Stoichiometry and the Mole

- 1 mol = 6×10^{23} atoms/molecules
- Limiting reactant: The reactant that is completely used up.
- Excess reactant: The reactant that has excess left after the chemical reaction.
- Molar gas volume is 1 mol = 24dm^3 at rtp, 22dm^3 according to Ideal Gas Law
- Molar mass = sum of relative atomic mass in particle / mol
 - o Note: rate of diffusion is dependent on M_r
- % purity = (Mass of pure substance / Mass of given sample) x 100%
- % yield = (Actual yield / Expected yield) x 100%
- Molarity = Amount (mol) / Volume (dm^3)

Acid Base Chemistry

Acid Base Reactions

- Acid + Base -> Salt + Water
- Acid + Carbonate -> Salt + Water + CO_2

- Acid + Metal \rightarrow Salt + H_2
- Acid + Sulfite \rightarrow Salt + Water + SO_2
- Base + Ammonium \rightarrow Salt + Water + Ammonia
- Alkali + Salt \rightarrow Salt + Insoluble Hydroxide (aka precipitation reaction)

Strength, Concentration and Basicity

- Strength = Degree of dissociation of an acid
- Concentration = Molarity of an acid
- Basicity is determined by number of H^+ ions: Monobasic, Dibasic and Tribasic

pH of a Solution

- $pH = -\log_{10}[H^+]$
- $[H^+] = 10^{-pH}$

Types of Oxides

- Basic Oxides (Most Metal) : react with acids to form salts
- Amphoteric oxides (Zn, Al, Pb) : react with both acids and bases to form salts
- Acidic Oxides (Nonmetal) : react with bases to form salts
- Neutral Oxides (Nonmetal) : do not react with acids or bases (CO , H_2O and NO)
- Hydroxides behave like metal oxides

Acid-Base Titration Curves

- Starting point at pH of acid/alkali
- Equivalence point is where all of the acid/base has been neutralised, take note of pH and Volume
- Equivalence point at 7 if strong acid strong base
- Equivalence point at 5 if strong acid weak base (estimate)
- Equivalence point at 9 if weak acid strong base (estimate)
- Ending point at pH of alkali/acid

Common Acids and Alkalis

Common Name	Chemical Name	Formula	Strength	Basicity
Hydrochloric Acid	Aqueous Hydrogen Chloride	HCl	Strong	Monobasic
Nitric Acid	Aqueous Hydrogen Nitrate	HNO_3	Strong	Monobasic
Sulfuric Acid	Aqueous Hydrogen Sulfate	H_2SO_4	Strong	Dibasic
Carbonic Acid	Aqueous Hydrogen Carbonate	H_2CO_3	Weak	Dibasic
Acetic Acid	Aqueous Ethanoic Acid	CH_3COOH	Weak	Monobasic

Caustic Soda	Sodium Hydroxide	NaOH	Strong	
Caustic Potash	Potassium Hydroxide	KOH	Strong	
Slaked Lime / Limewater	Calcium Hydroxide (slightly soluble)	Ca(OH) ₂	Strong	
Ammonia Solution	Aqueous Ammonia	NH ₃ (aq)	Weak	

Observations of Reactions

- Production of Gas
- Production of ppt
- Dissolution of a solid
- Colour change

Rules of Solubility in ppt reactions

- Most Chlorides, Iodides and Bromides soluble except for Lead (II) and Silver
- Most Sulfates soluble except for Barium, Lead (II) and Calcium
- Most Carbonates insoluble except for Ammonium and Group I
- Most Oxides and Hydroxides insoluble except for Group I. Calcium is sparingly soluble
- All Group I, Ammonium and Nitrates are soluble.
- Add Nitrate to dissolve
- Add Lead (II) Nitrate to precipitate

Ionic Equations

- Equation that shows exactly what particles take part in a reaction
- Eliminate ions which do not change in charge in entirety of reaction or are dissociated or reassociated
- All neutralisation reactions are: $H^+ + OH^- \rightarrow H_2O$
- Precipitation reactions only involve precipitate: $Ag^+ + Cl^- \rightarrow AgCl$
- Acid and metal reactions involve metal and H₂: $Zn + 2H^+ \rightarrow Zn^{2+} + H_2$
- Acid and soluble carbonate involve H₂O and CO₂: $2H^+ + CO_3^{2-} \rightarrow H_2O + CO_2$
- Displacement reaction involves displaced ions: $Zn^{2+} + Cu \rightarrow Zn + Cu^{2+}$

Redox

Oxidation States

- Oxidation states are the relative number of electrons an atom will inherit upon all molecules splitting apart
- Uncombined elements have oxidation number of 0
- Molecules (not ions) have total oxidation number of 0
- Ions have oxidation number of their total charge

- H is always +1, except in metal hydrides where it is -1
- O is always -2, except in peroxides where it is -1
- Group I elements are charge +1, Group II +2 and Group 3 +3
- Group VII elements are usually -1

Redox Reactions

- Reactions where one atom is reduced and one atom is oxidised
- **Oxidation is loss of electrons OIL**
- **Reduction is gain of electrons RIG**
- Oxidation state of [element] [increased/decreased] from [original oxidation state] in [reactant] to [new oxidation state] in [product], hence [reactant] was [oxidised/reduced]
- The [reducing/oxidising] agent is [substance] because it [donated/accepted] electrons [to/from] [other substance] to form [product of other substance], causing [other substance] to be [reduced/oxidised]
- All reactions which involve changing element molecules to compounds or vice versa are redox

Oxidising agents and reducing agents

- Take the case where X and Y reacted and X gained electrons from Y
- X is reduced by Y, hence X is the oxidising agent
- Y is oxidised by X, hence Y is the reducing agent
- To test for oxidising agent, use aqueous KI, positive result is when solution turns brown/yellow (I^- oxidised to I_2)
- To test for reducing agent, use acidified $KMnO_4$, positive result is when solution turns colourless (MnO_4^- reduced to Mn^{2+})
- Strongest oxidising/reducing agent is determined by how many other particles it can oxidise/reduce

Chemical Energetics / Enthalpy

Energy conversion in reactions

- **Heat or light change**
- Electrical change
- Nuclear reactions (fission or fusion)

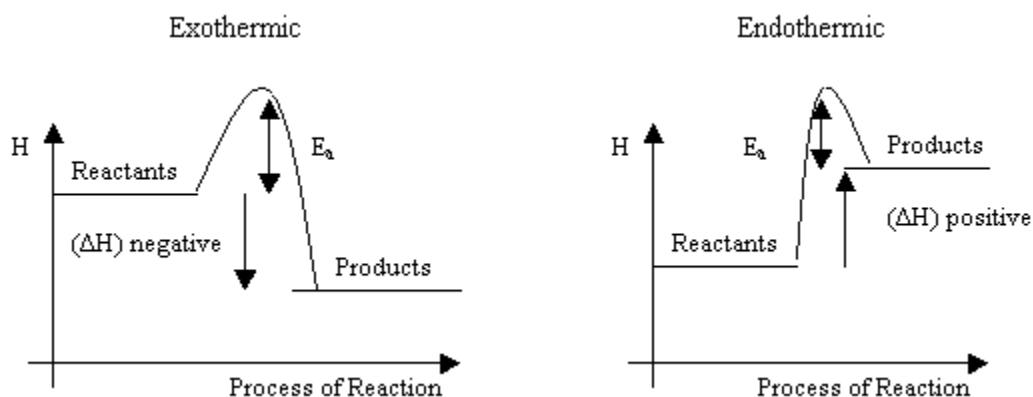
Exothermic and Endothermic reactions

Exothermic (out-heat)	Endothermic (in-heat)
Feels hot	Feels cold
Energy released (atoms need energy to prevent from being bonded)	Energy taken in

Bond forming (Liquid to solid is exothermic)	Bond breaking
Energy of reactants > Energy of products	Energy of reactants < Energy of products
ΔH negative	ΔH positive
Combustion, respiration, neutralisation, freezing, condensation and dilution of concentrated sulfuric acid	Thermal decomposition, melting and boiling

ΔH = energy from bond breaking - energy of bond making
= bond energy of products - bond energy of reactants

Enthalpy Diagrams / Energy Profile Diagrams



Reaction Kinetics

Factors affecting rate of reaction

1. Number of collisions (aka opportunities to react)
2. Activation energy E_A i.e. the minimum energy level for reaction to occur

Methods of increasing rate of reaction

- Increase concentration (1)
 - More particles per unit volume -> frequency of collisions ++
- Increase pressure for gas (1)
 - Particles in closer proximity -> frequency of collisions ++
- Increase temperature (1 and 2)
 - Increase in particles having activation energy / particles collide harder -> more successful reactions
 - Particles move faster -> frequency of collisions ++
- Introduce catalyst (2)

- Reduce in required activation energy -> more successful reactions
- Increase surface area (1)
 - Area of contact increases -> frequency of collisions ++

Catalysts

- Component in a reaction which speeds up the rate of a reaction
- Catalysts are neither consumed nor produced in the reaction, they only take part
- Catalysts provide an alternate reaction pathway to reduce the activation energy of a certain reaction, hence increasing the speed of reaction.
- Include enzymes as biological catalysts

Reaction Rate Graphs

- Graphs which have data which change according to rate of reaction/progress of reaction
- Measures data to determine rate of reaction
 - Amount of reactant remaining e.g. titrating samples taken at time intervals
 - Color intensity
 - Mass and pressure
 - Temperature

Periodic Table

Structure of the Periodic Table

- Way of arranging and classifying elements, arranged by proton number
- Vertical columns are called groups
 - Numbered in Roman Numerals
 - Elements in same group typically have same number of valence electrons
 - Elements in same group form ions of similar charge
 - Elements in same group tend to have similar chemical properties
- Horizontal rows are called periods
 - Progress from metals to non-metals when going left to right

Metals vs Non-metals

Metals	Non-metals
<ul style="list-style-type: none"> ● Solid at room temperature except Hg, some group I metals ● High mp and bp except Group I ● Good conductors of heat and electricity ● Shiny, ductile, malleable and have tensile strength ● Basic or Amphoteric oxides ● Form positive ions / cations 	<ul style="list-style-type: none"> ● Often gases except C, P, S and Si ● Low mp and bp for small molecules ● High mp and bp for macromolecules, e.g. C and Si ● Poor conductor of heat and electricity except graphite ● Dull and soft ● Acidic or Neutral oxides ● Form negative ions / anions

Trends in the Periodic Table

- Nuclear Charge
 - Amount of positive charge in the nucleus of an atom
 - Proton number of the atom
- Shielding electrons
 - All electrons not in valence shell
 - Dampens effect of nucleus on valence electrons
 - He has 0 shielding electrons
 - Remains the same across a period
 - Increase down a group
- Atomic Radii
 - Distance between outermost shell of electrons and center of nucleus of an atom
 - Determined by effect of nuclear charge on shielding electrons
 - Decreases down a period
 - Nuclear charge increases while shielding electrons remain constant
 - Valence electrons are increasingly pulled towards nucleus
 - Increases down a group
 - More electron shells are added as we go down the group
 - Valence electrons progress further from nucleus
- Ionization Energies
 - Amount of energy to remove one electron from each atom in one mole of gaseous element
 - Determined by effect of nuclear charge on valence electrons
 - Increases down a period
 - Nuclear charge increases while shield electrons remain constant
 - Force of attraction between valence and nucleus increases
 - Decreases down a group
 - More electron shells are added as we go down the group
 - Weaker force of attraction between further valence shell and nucleus
- Melting Points
 - Decrease down a group of metals
 - Force of attraction is between positive nuclei and delocalised electrons
 - Metals with larger atoms have delocalised electrons further from positive nuclei
 - Increase down a group of non-metals
 - Larger surface area for intermolecular forces with larger particles

Trends within Groups

- Group I Elements
 - Called Alkali metals
 - Density increases down group
 - Melting point decreases down group
 - Reactivity increases down group

- Form alkaline solutions and hydrogen gas with water
- Group VII Elements
 - Called Halogens / Halogen gases
 - Melting point increases down group
 - Reactivity decreases down group
 - More reactive halogen displaces less reactive halogen from its salt solution
 - More reactive particle creates a stronger / more stable / high-energy bond
 - Cl displaces Br and I salts
 - Br displaces I salts
 - Reacts with group I metals to form salts
- Group 0 Elements
 - Called Noble gases
 - Monatomic gases made of single atoms
 - Generally unreactive due to full valence shell of electrons
 - Density increases down group
- Transition Elements
 - Have high melting and boiling points
 - Have high density
 - Show variable oxidation states
 - Form colored compounds
 - Used as catalysts (Ni for margarine, Fe for Haber, V₂O₅ for Contact)

Group I vs Transition Metals

Physical Similarities
<ul style="list-style-type: none"> ● Shiny, malleable and electrically conductive
Chemical Similarities
<ul style="list-style-type: none"> ● Reacts with acids to form salt and hydrogen ● React with non-metals to form ionic compounds ● Burn in oxygen to form basic or amphoteric oxides

Physical Property	Group I Metals	Transition Metals
Outermost Orbital	s-orbital	d-orbital
Atomic Radius	Smaller	Larger
Melting/Boiling Point	Low	High
Density	Less Dense	Denser

Hardness	Malleable and Ductile	Hard
Chemical Property	Group I Metals	Transition Metals
Oxidation States	Only +1	Variable
Colour of Compounds	White solids Colorless solutions	Colored solids and solutions
Solubility of Compounds	All soluble	Not necessarily soluble
Catalytic Property	None	Present

- Physical patterns and interpretation
 - Group I soft, silvery and have low mp/bp
 - Transition metals hard and have high mp/bp
 - Similarities due to metallic bonding, differences due to atomic radii and valence shell
- Chemical patterns and interpretation
 - Group 1 are reactive to form white soluble salts
 - Transition metals form colored compounds. Transition metals can be used as catalysts
 - Group 1 only use outermost s-orbital for bonding
 - Transition metals use d-orbital for bonding

Qualitative Analysis

General Techniques in QA

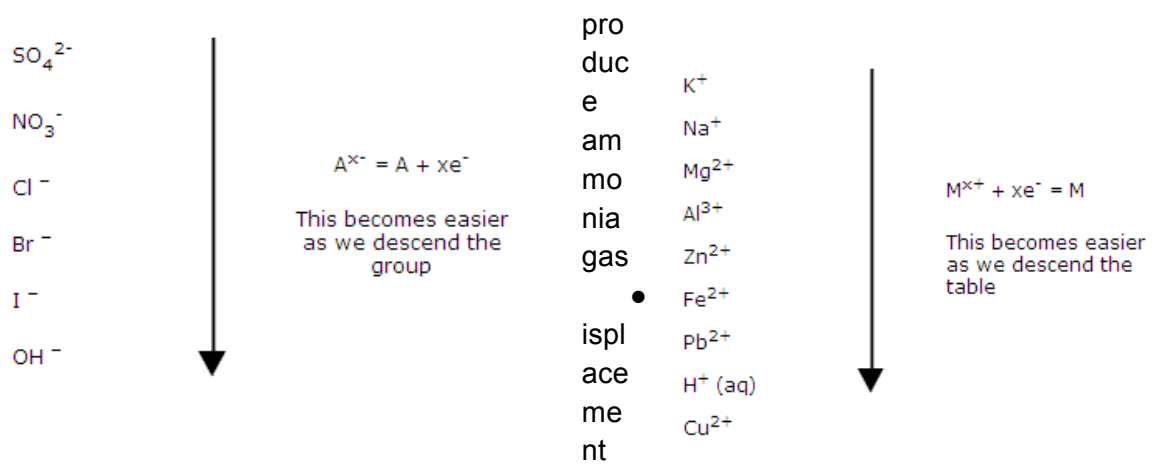
- Quantity of substance to be tested
 - Use approximately than 0.5cm^2 of solid (hemisphere of test tube)
 - Use no more than 2cm^2 of liquid
- Preparing solutions
 - Used deionised/distilled water
 - Avoid impurities in tap water
- Transferring chemicals
 - Use spatula to transfer crystals and powders, ensuring test tubes are dry but not hot
 - Check reagent label, remove stopper, pour reagent into test tube with rim touching test tube, replace stopper
 - Avoid spilling chemicals on hands, clothes or benches due to chemical stains or burns
- Adding a reagent
 - Add in small amounts first and observe, then shake and add more
 - Add to excess if needed
- Testing for Solubility

- Decant small amounts of ppt into separate test tube to test for solubility if possible
- Using the Bunsen Burner
 - Close air hole at beginning
 - Open gas tap and light bunsen burner
 - Open air hole to obtain non-luminous flame
 - Adjust gas tap to obtain desired strength of flame
- Heating Substances
 - When heating solid
 - Use test tube holder
 - Heat in dry test tube
 - Heat solid gently at first, observe for water vapour etc
 - If heating strongly, report melting
 - When heating liquid
 - Point mouth of test tube away from yourself and other people
 - Shake solution from time to time to let gases out
 - If heating gently, avoid boiling
- Testing for gases
 - Always test for gas when effervescence is observed
 - Record color, smell, reaction with litmus, splints and other reagents
- Recording Observations
 - Record observations immediately before proceeding to next test
 - Record all
 - Effervescence
 - Colour changes
 - Sound
 - Heat changes (feels warm / feels cold)
 - Precipitates and dissolutions
 - Negative results (no visible change is observed)
 - Uses names given by question, not by element
 - Do not conduct excess tests except for gases

Correct Observations	Incorrect Observations
White ppt in colourless solution	white solution is formed chalky suspension
black ppt in colourless solution	black solution
colourless solution	clear solution
ppt settles in solution	ppt settles to form a layer at the bottom with solution on top

Types of Reactions

- Formation of solids
 - Also known as precipitation
 - Formation of insoluble compound due to reaction
- Formation of water
 - Especially from acid-base reactions
 - $H^+ + OH^- \rightarrow H_2O$
- Formation of gas
 - Dilute acid reacts with carbonate to produce carbon dioxide
 - Dilute acid reacts with metal to produce hydrogen
 - Dilute acid reacts with sulphite to produce sulfur dioxide
 - Ammonium salt reacts with base to



- Only occurs if most reactive cation is not yet bonded to most reactive anion
- e.g. $NaNO_3$ and $AlCl_3$ will not displace but $NaCl$ and $Al(NO_3)_3$ will
- Combustion
- Metal and non-metal (formation of salt)
- Decomposition

QA of Anions

Reagent \ Anion	CO_3^{2-}	Cl^-	I^-	SO_4^{2-}
Ag^+	Ag_2CO_3 White ppt Soluble in acid Effervescence of CO_2	$AgCl$ White ppt	AgI Pale Yellow ppt	No ppt
Ba^{2+}	$BaCO_3$ White ppt Soluble in acid Effervescence of CO_2	No ppt	No ppt	$BaSO_4$ White ppt
Pb^{2+}	$PbCO_3$ White ppt	$PbCl_2$ White ppt	PbI_2 Yellow ppt	$PbSO_4$ White ppt

	Soluble in acid Effervescence of CO₂			
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- 2 whites, 1 color, 1 gasy and 1 unreactive
- To check for NO₃⁻, add NaOH (and Al(s)) and heat and look for ammonia gas

QA of Cations

Cation \ Reagent	NaOH	NH ₄ OH
Ca ²⁺	White ppt Insoluble in excess	No ppt
Zn ²⁺	White ppt Dissolves in excess Forms a colourless solution	White ppt Dissolves in excess Forms a colourless solution
Al ³⁺	White ppt Dissolves in excess Forms a colourless solution	White ppt Insoluble in excess
Pb ²⁺	White ppt Dissolves in excess Forms a colourless solution	White ppt Insoluble in excess
Cu ²⁺	Blue ppt Insoluble in excess	Blue ppt Dissolves in excess Forms dark blue solution
Fe ³⁺	Red-Brown ppt Insoluble in excess	Red-Brown ppt Insoluble in excess
Fe ²⁺	Green ppt Insoluble in excess Green ppt turns brown on standing	Green ppt Insoluble in excess Green ppt turns brown on standing

Cation \ Reagent	KI
Pb ²⁺	Yellow ppt
Al ³⁺	No ppt

- 4 whites, 3 colors and 1 unreactive
- Remember that ZAP of amphoteric oxides (and hence hydroxides) produce white ppt
- To check for NH₄⁺, add NaOH, heat gently and test for ammonia gas

QA of Gases

- To check for NH₃, let red litmus turn blue

- To check for CO₂, bubble through CaCO₃ solution and look for white ppt
- To check for Cl₂, let (blue) litmus bleach
- To check for H₂, add lighted splint and hear for 'pop' sound
- To check for O₂, add glowing splint and look for relight
- To check for SO₂, turns (acidified) purple KMnO₄ (VII) colourless

Chemical Equilibrium

Reversible Reactions

- Some reactions may be able to proceed forward and backward
- These reactions are called reversible reactions
- Signified by a double arrow between reactants and products
- Left to right is called forward reaction
- Right to left is called backward/reverse reaction
- Reversible reactions typically do not complete, hence a mixture of reactants and products are obtained
- When concentration of reactants and products is constant, mixture is at **equilibrium**



Dynamic Equilibrium

- Reactants decrease as forward reaction proceeds, producing more product
- Products decrease as reverse reaction speeds up, producing more reactant
- When rate of two reactions become equal, there will be no change in amounts and hence equilibrium
- When forward reaction takes place as fast as reverse reaction, equilibrium is achieved
- BUT reactions do not stop, hence it is considered a dynamic equilibrium

Le Chatelier's Principle

- Systems in equilibrium, when disturbed, respond by taking action to counteract the change
- If [factor] is changed, according to Le Chatelier's Principle, the system counteracts this [change in factor] by [counter to change in factor], thus favouring the [forward/reverse] reaction, thus the position of equilibrium shifts to the [right/left] and the yield of [product] [change in product].
- Typically encompasses change in
 - Concentration
 - Concentration changes and rate of reaction changes accordingly
 - System tries to maintain equilibrium of reactant and product by encouraging one type of reaction for its products
 - Equilibrium shifts to right when reactant is increased
 - Equilibrium shifts to the left when product is increased
 - Temperature

- Both rates of reaction change, but one changes more than the other -> disruption of equilibrium
 - System tries to maintain constant heat by encouraging one type of reaction for its enthalpy change
 - Equilibrium shifts to exothermic side when heat is removed
 - Equilibrium shifts to endothermic side when heat is introduced
 - Pressure (for gases)
 - System tries to maintain constant pressure by increasing or decreasing the amount of gas in a system
 - Equilibrium shifts to side with more moles of substance when pressure is decreased
 - Equilibrium shifts to side with less moles of substance when pressure is increased
- Recall other factors in reaction kinetics
 - Catalysts do not disturb the equilibrium since they accelerate both forward and reverse reactions

Metal Reactivity

Basic Reactivity

- How reactive a metal is is determined by how easily it loses its electrons
- More reactive metals lose electrons easier than less reactive metals
- Reactivity increases down a group, decreases down period
 - Outer shell electrons move further away from nucleus
 - Effect of nuclear charge on electrons are less
- Order can be determined by ranking the strength of reactions
 - Stronger reactions / more reactions of a chemical imply high reactivity

Reactions of Metals and Thermal Stability in the Reactivity Series

- Remember using mnemonics
 - People's Service Commission PSC
 - MAZIT or MACZIT with carbon
 - Low Class SinGaporean LCSG or Low-High Class Singaporean LHCSG with hydrogen
- Oxides and Acids
 - PSC form hydroxides, Magnesium sparingly reacts
 - MAZIT form oxides
 - L react with acids
 - CSG do not react with acid
- Strength as reducing agent decreases as reactivity decreases
 - Metals are less likely to give up electrons and oxidise themselves
 - K, Na are explosive with water while Ag and Au are inert

- More reactive metal will displace a less reactive metal from its compound
 - More reactive metal more ready to give up electrons and form a stronger chemical bond
 - A less reactive metal compound can be purified by smelting with a more reactive metal, more reactive metal becomes compound and less reactive is purified
 - Any oxide less reactive than carbon can be reduced and purified by the use of coke (C)
- Heating of Pure Metals, Oxides, Hydroxides, Carbonates and Nitrates

Heating of:	Pure Metal	Oxides	Hydroxides	Carbonates	Nitrates
PS	Metal Oxide and Peroxide	Stable	Stable	Stable	Metal Nitrite + O ₂
CMAZIT	Oxide (dec vigour)	Stable	Metal Oxide + H ₂ O	Metal Oxide + CO ₂	Metal Oxide + NO ₂ + O ₂
LC	Surface Oxide	Stable	Metal Oxide + H ₂ O	Metal Oxide + CO ₂	Metal Oxide + NO ₂ + O ₂
SG	Do not exist	Metal + O ₂	Do not exist	Metal + CO ₂ + O ₂	Metal + NO ₂ + O ₂

- All metal react with Cl₂ on heating to form chloride w/ decreasing vigour
- More reactive metal forms more thermally stable compound
 - Less reactive materials occur naturally since they do not form compounds easily
 - Less reactive materials hence are easier to obtain

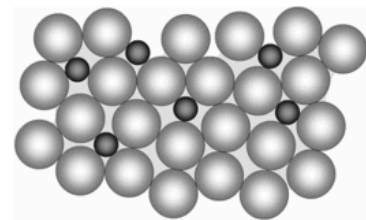
Corrosion of Metals

- Most metals usually are shiny and grey (Cu is red-brown)
- Metals naturally react with surroundings to form compounds, destroying the metal
- Process is called corrosion
- Rusting is the corrosion of iron in the presence of water and oxygen
 - Reduced by oxygen to form iron (II) ions
 - Surrounding water forms hydroxide ions
 - Hydroxide further oxidised iron to form iron (III) hydroxide
 - Eventually turning into hydrated iron (III) oxide or rust
- Conditions speeding up rusting
 - Accelerated when water has other ionic substances, increasing concentration of hydroxide ions
 - Salt water rusts iron faster than distilled water
 - Accelerated when in contact with less reactive metal, where iron will sacrifice electrons to the other metal
- Preventing rust (and other forms of corrosion)
 - Applying protective layer of paint, grease or metal (less reactive than iron) plating

- Risks scratching and exposing iron
- Sacrificial corrosion / galvanizing where another (more reactive) metal is in contact with iron to donate electrons and let the other metal corrode preferentially
 - Underwater pipes usually have magnesium blocks attached
- Corrosion is not universal
 - Aluminum Oxide is impervious, which does not allow oxygen to pass through the non-porous layer of oxide and hence protecting the inner layers of aluminum
 - Al hence seems a lot less reactive than other metals, despite its location in the reactivity series
 - Oxide layer can be chemically enlarged via anodising (introduction of a dye which can penetrate oxide layer), increasing the size of impervious layer

Metal Alloys

- Pure metals are soft and malleable, since rows of metal nuclei can slide easily
- Adding impurities in the form of metal or carbon strengthens and hardens a metal
- Different rows of metal nuclei are unable to slide due to impurities and uneven nuclei size
- Produced by smelting components together
- Most important is steel - Iron and Carbon, sometimes with other metals
 - Carbon -> Hardness and brittleness ++
 - Cobalt -> Magnetism ++
 - Tungsten -> Heat resistance ++
 - Manganese -> Hardness ++ (rock drills)
 - Silicon -> Springiness ++
 - Molybdenum -> Acid resistance ++



Alloy	Composition	Advantages	Uses
Mild Steel	99 - 99.5 Fe 0.15 - 0.25 C	Harder, stronger, more malleable and ductile than iron	Car bodies Reinforcement beams
Stainless Steel	90 - 95 Fe 5 - 10 Cr and Ni some C	Harder and more resistant to corrosion Shiny appearance	Cutlery Surgical instruments

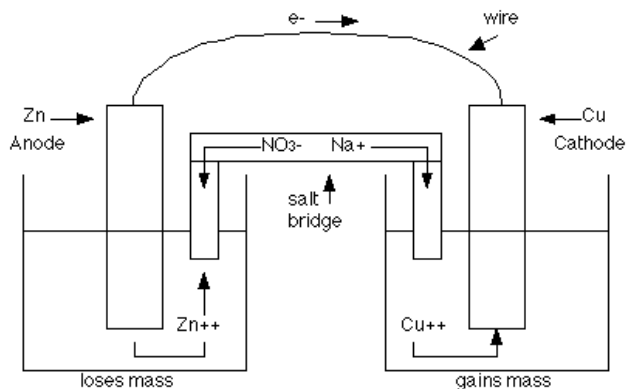
Metal Recycling

- Metals are a finite resource and can run out
- Recycling metals can slowdown the exhaustion of metal reserves
- Generally easy to recycle since can be smelted down
- Most recycled is Steel (Iron w/ <2% Carbon) and Aluminum
- Benefits:
 - Conserving metal reserves
 - Reduce cost of extracting new metals

- Reduce environmental pollution due to polluting mining processes
- Reduce load on landfills
- Disadvantages:
 - Possible high cost due to collection, transport, sorting, purification etc
 - Can produce pollutants
 - Does not apply to all alloys, some are inseparable

Electric Cells

- Arrangement to **produce** electron flow over a wire
- Consists of 2 electrodes immersed in an electrolyte and wire connecting the electrodes
 - Electrodes are different metals, hence have different reactivity
 - Electrolyte (solution with free ions) provides free, mobile ions to conduct electric charge between the electrodes. Usually contains metal ions of the less reactive metal.
 - Wire allows electrons to pass through to the other electrode
- Potential difference is generated between electrodes due to different reactivity, more reactive metal tries to donate electrons to less reactive metal



An electrochemical cell of the reaction:
 $Zn(s) + Cu^{++}(aq) \rightarrow Zn^{++}(aq) + Cu(s)$

Metal	More reactive	Less reactive
Electric Terminal	- Negative (donates e^-)	+ Positive (receives e^-)
Chemical Process	Oxidation / loss of e^-	Reduction / gain of e^-
Electrode Name Determined by redox type	Anode (A node has O Xidation)	Cathode (RED uction at CAT hode)
Movement of metal nuclei	Electrode to electrolyte	Electrolyte to electrode
Chemical Equation	$M(s) \rightarrow M^+(aq) + e^-$	$N^+(aq) + e^- \rightarrow N(s)$
Flow of electrons	Electrode to wire/load	Wire/load to electrode

- Where M is cathode material and N is electrolyte cation
- Overall equation $M + N^+ \rightarrow M^+ + N$, with electric flow through wire

- As the reaction progresses, more reactive metal becomes smaller and smaller while cations inside electrolyte are deposited on the cathode and the concentration of cation in the electrolyte decreases (Cu^{2+} electrolytes lose their color)
- Least reactive cation is **deposited at cathode**
 - Most of the time will be H^+ deposited to give H_2 gas
- The larger the difference of reactivity between metals, the larger the voltage
- To use separate electrolytes, two electrolytes are joined with a salt bridge which has electrolytes of good conductivity inside to allow charge flow between electrolytes

Fuel Cell

- Prime application of electric cell is the hydrogen fuel cell
- Hydrogen as a fuel
 - Burns cleanly to produce H_2O , instead of petrol which produces CO and CO_2
 - Produces twice as much energy as common fuels
 - Cannot be transferred as a liquid
- Fuel cells are used to react hydrogen with oxygen and get electrical energy directly
 - H_2 gas in contact with anode, dissociates to H^+ and e^-
 - $\text{H}_2 \rightarrow 2\text{H}^+ + 2\text{e}^-$
 - H^+ passes through anode and electrolyte to cathode
 - e^- passes through terminals and load to cathode
 - Meets with O_2 at cathode to form water
 - $4\text{H}^+ + 4\text{e}^- + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
 - Overall equation $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$
- Advantages of fuel cell:
 - Supplies electricity like battery, can easily control voltage etc
 - Electrodes do not deteriorate, can supply energy as long as fuel is supplied
 - Can operate in silence due to lack of moving parts
 - Efficient, produces energy at 50% efficiency vs 20 - 30% for petrol
 - Generate energy electrochemically hence consume fuel only as much energy as the load draws
 - Produce water, hence no pollutant emissions

Electrolysis

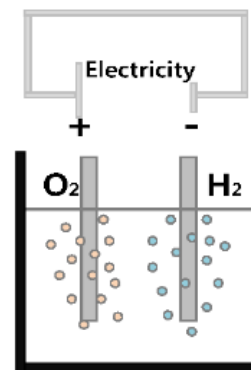
Electrolysis

- Fuel cells generate electricity using chemical change
- Electrolysis causes chemical change using electricity
- Electrolyte: A substance which can only conduct in the molten or aqueous form
- Electrolysis: the conduction of electricity by an ionic compound, leading to the decomposition of the electrolyte
- When an electric current is passed through an electrolyte, the positive ions are attracted to the cathode while the negative ions are attracted to the anode

- As a result, electrolyte is discharged (lose ion status) and certain products are formed at each electrode

Electrolytic Cells

- Arrangement to use electric flow to **decompose** an electrolyte
- Consists of 2 electrodes immersed in an electrolyte with power supplied to electrodes
 - Electric current is supplied to electrodes
 - Electrodes enact charge given by electric charge by attracting specific ions in the electrolyte
 - Flow of free mobile ions across electrolyte result in decomposition of electrolyte near the electrodes



Electric Terminal (swapped)	+ Positive (receives e^-) (ANode)	- Negative (donates e^-)
Chemical Process	Oxidation / loss of e^-	Reduction / gain of e^-
Electrode Name Determined by redox type	Anode (ANode has OXidation)	Cathode (REDuction at CAThode)
Movement of metal nuclei	Electrode to electrolyte	Electrolyte to electrode
Chemical Equation	$M^+ (aq) \rightarrow M + e^-$	$N^+ (aq) + e^- \rightarrow N$

Flow of electrons	Electrode to source	Source to electrode
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- Assuming electrolyte MN
 - N^- and OH^- migrates to anode, substances produced leaving behind H^+
 - M^+ and H^+ migrates to cathode, substances produced leaving behind OH^-

- Check electrochemical series for which is discharged

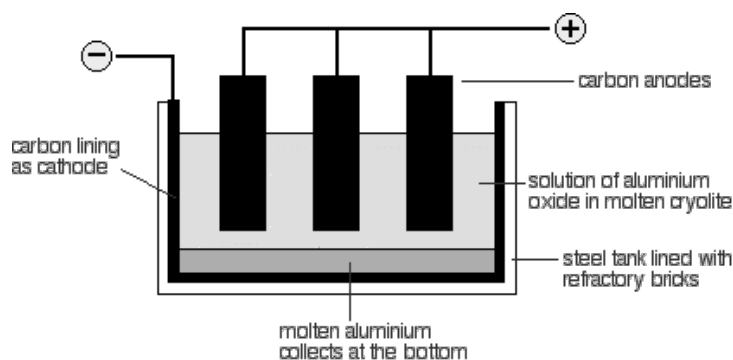
Factors affecting Electrolysis

- Different electrolysis processes result in the formation of different substances at the electrodes, and are affected by:
- Preferential Discharge of Ions
 - Some ions are more reactive than others, hence some are easier discharged than others.
 - Most easily discharged ions which have migrated to the electrode will be preferentially discharged, refer to electrochemical series
 - Cation series is reactivity series with Calcium and Sodium swapped
 - Based on the electrochemical series, [ion] [gains/loses] electrons more readily than [other ion], hence [ion] is preferentially reduced/oxidised
- Concentration Effect
 - Increasing the concentration of an ion promotes its discharge
 - Order of discharge according to preferential discharge can be modified
 - The concentration of [ion] is high enough for the concentration effect to take place, due to the closeness of the positions of [ion] and [other ion] in the electrochemical series. Thus, [ion] is preferentially oxidised. (Cations typically are not affected by concentration effect)
 - Though the concentration of [ion] is high, the distance between the positions of [ion] and [other ion] in the electrochemical series is too great, thus [ion] is still preferentially reduced/oxidised
- Electrode Material
 - Electrodes such as carbon and platinum are inert electrodes, which take no part in electrolysis.
 - Carbon is attacked by oxygen
 - Platinum is attacked by chlorine
 - Some electrodes can influence the ionic discharge due to participation in the redox process and reduce the amount of energy needed to discharge a certain ion, hence they are called active electrodes
 - Additionally, some electrodes require less energy to ionise themselves than to discharge the electrolyte, resulting in the preferential charge of some types of electrodes instead of discharge of the electrolyte.

Electrolytic Purification

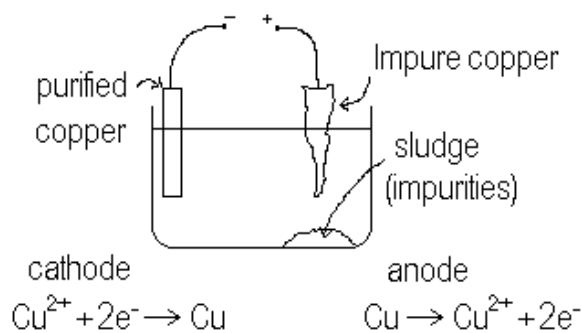
- Electrolysis can be used to purify metals, producing pure metal at one of the electrodes

Aluminum Purification



- Bauxite (Al_2O_3) dissolved in molten cryolite is used as electrolyte, graphite used as electrodes
 - Cryolite is used to reduce energy costs and dissolve bauxite instead of having to melt bauxite at 2000°C
 - Graphite used as inert electrode
- Al^{3+} (aq) migrates to cathode and is discharged, hence purified into Al (l)
- O^- (aq) migrates to anode and is discharged, hence forming O_2 (g)
 - O_2 is a by-product, if using platinum electrodes can be used to produce pure O_2 to sell
- Molten aluminum is obtained

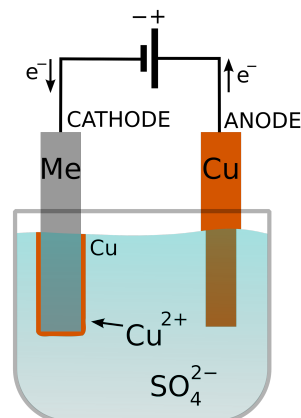
Copper Purification



- Copper (II) sulfate is used as electrolyte, pure copper is used as cathode and impure copper is used as anode
 - Impure copper charges and dissolve rather than oxidising electrolyte
 - Cathode reduces copper ions before hydrogen ions, depositing copper
 - Impurities are left as ppt in the electrolyte
- Cu (s) from the impure copper is charged and dissolved to form Cu^{2+} (aq) ions
- Cu^{2+} (aq) migrates to the cathode and is discharged, hence forming Cu (s)
- Anode reduces in size
- Copper (II) sulfate remains in concentration
- Purified copper enlarges

Electroplating

- Process to coat the surface of one metal with another metal
 - For appearance, decoration or for protection against corrosion
- Considering copper plating, copper (II) sulfate is electrolyte, object to be plated is the cathode while the anode is a piece of pure copper
 - Copper charges and dissolve rather than oxidising electrolyte
 - Pure copper reduces copper ions before hydrogen ions
- Cu (s) from the impure copper is charged and dissolved to form Cu^{2+} (aq) ions
- Cu^{2+} (aq) migrates to the cathode and is discharged, hence forming Cu (s)
- Anode reduces in size
- Copper (II) sulfate remains in concentration
- Cathode is copper plated



Organic Chemistry

Organic Molecules

- Carbon has 4 valence electrons and hence easily form covalent bonds with other atoms
- Any molecule with carbon is considered to be an organic molecule
- Typical organic molecule consists of carbon chain backbone with functional groups added in with all extra unbonded valence electrons bonded to hydrogen
- Chemical properties:
 - Nonpolar and do not ionise in solution
 - Insoluble in water unless containing polar group (O-H/F-H/N-H)
 - Mix well with nonpolar solvents
 - High melting and boiling points
 - Thermally unstable; flammable and decompose exothermically into carbon dioxide and water readily

Naming Conventions of Organic Molecules

- All organic molecules have a name which conveniently describe its structure
- First consideration is length of longest carbon chain which determines base prefix

Carbons	1	2	3	4	5	6
Prefix	Meth-	Eth-	Prop-	But-	Pent-	Hex-

- Next consideration is the dominant functional group(s) and its positions in the carbon chain which determine base suffix and locator numerals
- Determine which end to start counting carbon chain by ensuring dominant functional group is bonded to smallest carbon number

Functional Group	None Alkane	C=C Bond Alkene	C≡C Bond Alkyne	C-O-H Bond Alcohol	C=O, C-O-H Bond Carboxylic Acid
Suffix	-ane	-ene	-yne	-anol	-anoic acid

- Determine additional functional groups and positions
 - Arrange additional functional group prefixes in alphabetical order
 - e.g. **B**romo**c**hlorobutane
- Add di/tri/etc prefixes to functional group prefixes/suffixes to indicate multiple functional groups
 - di-bromo (two Br) / di-ene (two C=C bonds)
- Add number of carbon attached to for functional groups not on Carbon 1
 - e.g. 2,3 dibromo prop-2-ene

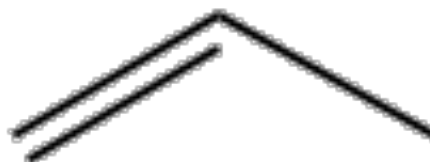
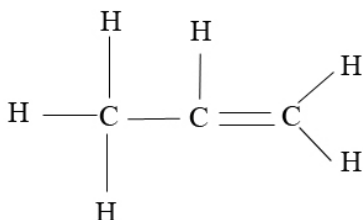
Organic Properties

- Homologous Series
 - Family of compounds which follow a structural pattern
 - Follow a general structural formula
- Functional group
 - Series of atoms or bond which determine a chemical property
- Saturated molecules
 - Only consist of single C-C bonds
 - No C=C and C≡C bonds
 - "Saturated" and has most number of C-C single bonds

Molecule Formulae

- Empirical formula
 - Simplest formula of a compound
 - Displays ratio of atoms which a molecule comprises
 - e.g. Propene empirical formula is CH₂ since C and H atoms are in a 1:2 ratio
- Molecular formula
 - Indicates actual numbers of each atom in a molecule of a substance
 - Each element represented once
 - Organised according to C_xH_yO_z
 - e.g. Propene molecular formula is C₃H₆ since there are 3 C and 6 H atoms
- Full molecular formula
 - Indicates which atoms/functional groups are bonded to what carbons
 - Elements represented multiple times
 - e.g. Propene full molecular formula is CH₂CH₁CH₃
- Full structural formula

- Also called display formula
- Visual representation of which atoms are bonded to what with what kind of bonds
- Bonds usually drawn at right angles
- e.g. Propene structural formula is (below left)



- Line formula
 - ~~Simpler version of full structural formula~~
 - ~~Drawn in hexagonal template to explicitly state carbons~~
 - ~~Each unmarked hexagonal vertice implies carbon~~
 - ~~Functional groups added as marked vertices~~
 - ~~All other bondable zones assumed to contain hydrogen~~
 - e.g. Propene line formula is (above right)

Functional Groups

- Common atomic structures found within organic compounds
- Contribute to chemical and physical properties of organic molecule

Alkanes

- “Basic” organic molecule, all organic molecules are derived from this base unit
- Suffix -ane
- Characteristic structure consists of chain of carbons bonded by single bonds only
 - Can be called “saturated” carbon chain because only C-C bonds are present
- Mostly liquid at room temperature, except first 4 (1 to 4 carbon) which are gases
- Chemically inert except for combustion (alkane → H₂O and CO₂) and substitution with chlorine (alkane + Cl₂ → chloroalkane + HCl)

Alkenes

- Characteristic functional group structure is the double C=C bond
 - Since there cannot be a double bond between one single C atom, methene is not a legitimate molecule
- Suffix -ene, -diene, -triene etc
- General formula for alkenes with one double bond is C_nH_{2n}
- Most liquid or gases at room temperature

- More reactive than alkanes, C=C bond allows it to undergo addition reactions where multiple molecules are combined to one molecule
 - Combustion (alkene \rightarrow H₂O and CO₂)
 - Hydrogenation (alkene + H₂ \rightarrow alkane)
 - Hydration (alkene + H₂O \rightarrow alcohol)
 - Bromination (alkene + Br₂ \rightarrow Dibromoalkane)
 - Can also be used to test for unsaturation
 - Brown solution turns colourless means solution contains unsaturated organic molecules
 - Note that Br atoms result in adjacent C atoms, not same C atom or non-adjacent C atom (e.g. 2,3 dibromo is okay, 2,4 dibromo is not)
 - Polymerisation (alkene + alkene \rightarrow polyalkene)

Alkynes

- ~~Characteristic functional group structure is the C≡C triple bond~~
 - ~~Since there cannot be a triple bond between one single C atom, methene is not a legitimate molecule~~
- ~~Suffix -yne, diyne, triyne~~
- ~~General formula for alkynes with one triple bond is C_nH_{2n-2}~~

Alcohol

- Characteristic functional group structure C-O-H hydroxyl group
- Suffix -ol, -diol, -triol
- General formula for alcohol with one hydroxyl group C_nH_{2n+1}OH
- Mostly liquid at room conditions and soluble in water
- Preparation
 - Alkene preparation
 - Alkenes mixed with water
 - Mixture of ethene and steam is passed over phosphoric acid (catalyst) at a high temperature of 300°C with 60 atm pressure
 - Fermentation
 - Produces ethanol from sugar
 - Sugar/Glucose -(heat, yeast, aqueous, anaerobic environment \rightarrow environment for zymase)-> 2 Ethanol + 2CO₂
- Reactions
 - Combustion (alcohol + O₂ \rightarrow H₂O and CO₂)
 - Dehydration (alcohol \rightarrow alkene + H₂O)
 - Heat alcohol in concentrated sulfuric acid (as dehydrating agent to absorb H₂O)
 - Oxidation (alcohol + O₂ \rightarrow carboxylic acid + H₂O)
 - Esterification (alcohol + acid \leftrightarrow ester + H₂O)
- Used as solvent (due to polar and nonpolar bonding sites), drink as well as fuel for vehicles

Carboxylic Acid

- Characteristic functional group structure C=O, C-O-H carbonyl group
- Suffix -anoic acid, -dianoic acid, -trianoic acid
- General formula for carboxylic acids with one carboxyl group is $C_nH_{2n+1}COOH$
- Mostly liquid at room conditions and soluble in water
- Do not dissociate completely in aqueous solution, weak acids
- Preparation
 - From alcohol (alcohol + $O_2 \rightarrow$ carboxylic acid + H_2O)
 - From methane
 - Methane produces methanol and CO, which are reacted with a catalyst to form ethanoic acid
- Reactions
 - Acidic properties
 - Reacts with base
 - Reacts with reactive metals
 - Reacts with carbonate
 - Ester formation (carboxylic acid + alcohol \rightarrow ester + H_2O)
- Used to make esters which are used for solvents and flavouring, insecticides and for drugs
 - Vinegar is aqueous ethanoic acid

Manufacturing Hydrocarbons

- Common hydrocarbons include crude oil, petroleum, natural gas and coal
- Specific hydrocarbons are obtained by fractional distillation of crude oil
- Different hydrocarbons with varying boiling points are separated
- Cracking
 - Large hydrocarbons can be obtained through fractional distillation
 - Cracking breaks down heavy fractions into lighter fractions which are in more demand
 - Purpose:
 - Turns long hydrocarbons into alkene and sometimes hydrogen, carbon and alkanes
 - Changes naphtha into alkenes
 - Creates hydrogen gas
 - Thermal cracking is where alkanes are heated to $800^\circ C$ in high pressure, hydrocarbons naturally break down
 - Catalytic cracking is where alkane molecules are passed over catalyst at $300^\circ C$, with catalysts such as alumina (Al_2O_3) and silica (SiO_2)
 - After cracking, products is separated by fractional distillation again

Combustion of Hydrocarbons

- Complete combustion is when hydrocarbon is burned to produce carbon dioxide and water only, occurs only in oxygen abundant environments

- Insufficient oxygen causes incomplete combustion to produce carbon monoxide and loose carbon/soot
- Amount of soot produced is dependent on % mass of carbon in a hydrocarbon

Reaction Flow

