KiMiYA
LEARNING PLACE

‘O’ Level Science Chemistry
Quick Revision Notes

- To check purity of a substance
  – Determine melting or boiling points (the only method)

Common Chemicals

Common Name	Chemical Name	Common Name	Chemical Name
Limestone	Calcium carbonate	Lime / quick lime	Calcium oxide
Limewater	Calcium hydroxide	Aqueous Ammonia	Ammonium hydroxide

Colour of some substances

Colour	Substances	Colour	Substances
White	Group I, II and III, ammonium compounds, Anhydrous $CuSO_4$	Black	CuO (s), I_2 (s)
		Shiny/grey	Most metals
Yellow	PbI, AgI, hot ZnO, Fe^{3+} (aq)	Blue	Cu^{2+} salts
Red-brown	Br_2(*l*), Copper Metal, $Fe(OH)_3$ (s)	Green	$Fe(OH)_2$ (s), some Cu^{2+} salts Fe^{2+} (aq), Cr^{3+}

1. Separation techniques

Method	Substance to separate	Substance(s) obtained	Concept behind method	Remarks
Filtration	▪ Insoluble solid from a liquid. ▪ Soluble solid & insoluble solid	Filtrate Residue (Solid)	▪ Different solubility in the liquid	▪ Large insoluble solid trapped by filter paper
Decanting	Large insoluble solid from liquid.	-	▪ Large insoluble solid	-
Evaporate to dryness	A soluble solid from a liquid.	Dry / anhydrous salt	▪ Different physical states	▪ Not for solids that decompose on heating
Crystallisation	▪ A pure solid from an impure solid. ▪ A soluble solid from a liquid. ▪ Substances which decompose.	Well-formed crystals of the pure solid (solute)	▪ As temperature decreases, **solubility** decreases. ▪ When saturated solution is cooled, crystals are formed.	▪ For solids that decompose on heating ▪ To retain water of crystallisation (hydrated salt)
Sublimation	A mixture of solids which one of them sublimes.	Pure solid which sublimes.	▪ One substance changes **directly** from solid to vapour on heating.	▪ E.g. iodine, moth balls, dry ice.
Simple distillation	A solution of a dissolved solid (solute) in a liquid.	Pure liquid as distillate.	▪ Different physical states	▪ Liquid is heated into vapour, cools and condenses into pure liquid in condenser and collected as **distillate**.
Fractional distillation	Mixture of miscible liquids	Pure liquids as distillate.	▪ Difference in boiling point	▪ Uses a fractionating column. ▪ Liquid with the lower bp will be distilled out first.
Separating funnel	Mixture of two immiscible liquids.	Two separated liquids	▪ Difference in densities	▪ Mixture separates into two layers
Paper chromatography	Solution containing mixture of small amounts of solids	Solutes separated in the chromatogram.	▪ Different solubility of solutes in the solvent used	▪ More soluble solutes travel a longer distance with solvent.

2. Particulate Nature of Matter

- Kinetic Particle Theory of Matter
  - Particles are too small to be seen directly
  - There are **spaces** between particles
  - Particles are in **constant random motion**; move in different speeds in different states
  - showed in Diffusion
    - *Rate of diffusion depends on **molecular mass** and **temperature***
- Particulate models of matter in different state
  - Describes the ***arrangement and motion of particles***

Physical State	Solid	Liquid	Gas
Arrangement	Orderly, closely packed	Disorderly, closely packed	Disorderly, far apart
Motion	Vibrate about fixed position	Slide pass one another	Free to move at high speeds in all direction

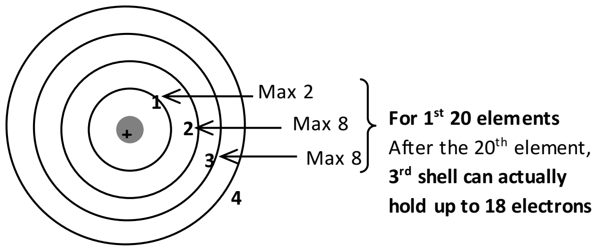
- **Temperature** of a matter is due to the **speed/kinetic energy of the particles** within it.
- Matter expands due to the **increase in distance between its particles**.
- Matter may be coloured but its **particles are not coloured**.

3. Element, Compound and Mixture

	Element	Compound	Mixture
Definition	▪ Pure substance ▪ Cannot be broken down into simpler substances by chemical means.	▪ Pure substance ▪ Contains two or more different elements chemically combined together.	▪ Contains two or more substances that are not chemically combined together.
Classification by...	▪ Physical state ▪ Periodic Table	▪ Types of bonding (Covalent / Ionic / Both)	▪ Mixture of elements ▪ Mixture of compounds ▪ Mixture of both elements and compounds.
It can exist as…	▪ Atoms (smallest unit) ▪ Molecules (in covalent bonding)	▪ Molecules (covalent compound) ▪ Ions (ionic compounds)	
Composition	▪ Fixed composition	▪ Fixed composition of the elements by mass.	▪ Variable composition.
Melting and boiling points	▪ Fixed	▪ Fixed	▪ Variable (A range of temperature)
Properties	▪ Own characteristic physical and chemical properties.	▪ Physical and chemical properties are different from its elements.	▪ Does not have its own properties. ▪ Has same properties as its components.
Formation	-	▪ Involves energy changes for chemical reaction to occur	▪ Mixing with no energy required
Separation	▪ Cannot be separated or broken down further.	▪ Cannot be separated by physical means. ▪ Requires energy (usually involves decomposition)	▪ Easily separated into its components by physical means without chemical reaction ▪ Involves little or no energy

4. Atomic structure

Subatomic particle	Symbol	Relative mass	Charge
Proton	p	1	1+
Neutron	n	1	0
Electron	e^-	$\frac{1}{1840}$	1-



- Atom
  - electrically neutral
  - equal no. of protons and electrons
- Isotope
  - atom of the same element but with different no. of neutrons
- Electronic configuration
  - shows the arrangement of electrons in an atom

***All atoms** wanted to obtain a **stable electronic configuration of a noble gas**.

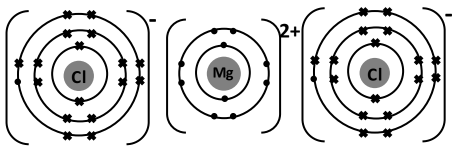
Therefore,

1) Atoms of the same element react to form molecules (by sharing valence electrons), **OR**
2) Different elements react to form stable compounds (by gain or lose valence electrons OR share valence electrons)

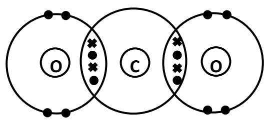
5. Chemical Bonding, Structure and Properties

	Ionic Bonding	Covalent Bonding
Definition	Strong electrostatic force of attraction between oppositely charged ions	Bond formed by the sharing of at least one pair of valence electrons between atoms
Formation of bond	▪ By **electron transfer** from metals to non-metals ▪ Metal atom loses e^- while non-metals gain e^-	▪ By **electron sharing** among atoms ▪ Non-metals share electrons ▪ No. of pairs of shared e^- = No. of covalent bonds
Description of particles formed	▪ Positive and negative ions ▪ **Simple ions/Polyatomic ions**	▪ Between atoms of same element ▪ Between atoms of different elements (compound)
Structure	<u>**Giant Ionic Lattice**</u> ▪ **Large number** of oppositely charged ions arranged in a repetitive, orderly manner, held in place by strong ionic bonds	<u>**Simple Molecular Structure**</u> ▪ Small molecules ▪ Strong covalent bond **within** molecule ▪ Weak intermolecular forces of attraction **between** molecules

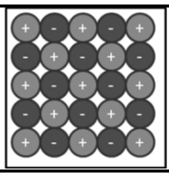
Dot and cross diagrams to show bonding



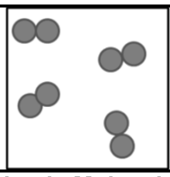
Ionic Bonding ($MgCl_2$)



Covalent bonding (CO_2)



Giant Ionic or Crystal Lattice



Simple Molecular Structure

Physical Properties of Different Structures

- Physical state / appearance depends on
  – type of bonding or attraction between particles in the substance
  – the arrangement of particles in the substance
- Melting and boiling point depends on
  – amount of energy required to overcome attraction
  – type and strength of bonding or attractive forces to be overcome / broken
- Volatility depends on boiling point
- Solubility depends on whether the particles can / cannot attract the solvent molecules
- Electrical conductivity
  – can conduct if there are mobile ions / electrons in the substance

	Giant Ionic Lattice	Simple Molecular Structure
Physical state	▪ Crystalline **solid** ▪ Flat sides and Regular shape ▪ Hard	▪ Most are **liquids or gases** at room temp
Melting and boiling points	**High m.p. and b.p.:** - Large amt of energy needed to overcome many strong electrostatic forces of attraction between oppositely charged ions	**Low m.p. and b.p.:** - Low energy needed to overcome weak intermolecular forces of attraction between molecules
Volatility	**Not volatile:** Strong ionic bonds hold the ions together	**Volatile** and evaporate easily to give a smell
Solubility	▪ **Soluble in water,** giant lattice breaks down and ions are mobile ▪ **Insoluble in organic solvent**	▪ Most are insoluble in water. ▪ Dissolves in organic solvent.
Electrical conductivity	▪ Cannot conduct in solid - ions are in fixed position ▪ **Can conduct in molten or aqueous** as ions are mobile	▪ **Do not conduct electricity at all** - molecules contain no mobile ions or e^-

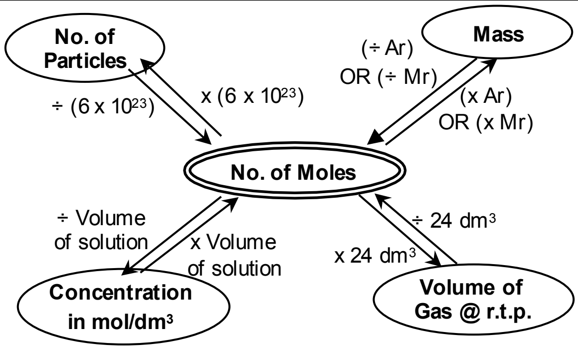
Note: Not all ionic compounds are soluble in water!!! e.g. $PbSO_4$

6. Mole Concept and Calculation

Relative atomic mass – **average mass** of one atom compared to the mass of $^{1}/_{12}$ of a carbon-12 atom
Relative molecular mass – average mass of one molecule compared to the mass of $^{1}/_{12}$ of a carbon-12 atom
One mole of a substance contains **6×10^{23}** particles. (Particles can be atoms, molecules, ions or even electrons)

- Molar mass – **mass of 1 mole** of any substance (in g/mol)
- Percentage composition – % by mass of each element present in a compound
- Avogadro's law – all gases with 1 mole of particles have equal volumes at the **same temperature and pressure**
  – Molar gas volume at r.t.p.= **24 dm³/mol**
- Concentration of Solutions – the amount of solute per unit volume of the solution
  – g/dm³ or **mol/dm³** or molarity (M)
- **Balanced** Chemical Equations
  – gives **molar ratio** or **volume ratio** of reactants and products
- **Limiting reactant** – limits the amount of products formed

1 mol of a substance: Contains 6×10^{23} particles
Has a mass of its Ar or Mr (in grams)
Occupies 24dm³ if it is a gas



Empirical formula – shows the ratio of the type of atoms present
Molecular formula – shows the actual number and type of atoms present
Structural formula – shows how atoms are joined in the molecules

7. Periodicity

Elements in same **Period** – same number of **shells**

Elements in same **Group** – same number of **valence electrons**

- Transition metals
  – variable oxidation states, form coloured compounds, used as catalyst
  – high mp and high densities

Group I	Group VII	Group 0
Shiny, silvery, soft Low density, low mp React easily in air. React vigorously with cold water.	Diatomic molecules Low mp and bp Very reactive non-metals More reactive halogen displaces less reactive halogen from its ions	Monoatomic Colourless gas at r.t.p Very low mp and bp Stable electronic configuration Unreactive
Li, Na, K, Rb, Cs ↓ ***Density increases*** ***Reactivity increases*** ***Melting point decreases***	F, Cl, Br, I, At ↓ ***Mp & Bp increases*** ***Becomes darker*** ***Change from gas to solid*** ***Reactivity decreases***	He, Ne, Ar, Kr, Xe, Rn ↓ ***Density increases*** ***Melting point increases***

Group I: Down the group,

valence electron further away from positive nucleus

$\Downarrow$

attraction is weaker $\Rightarrow$ easier to lose valence electron $\Rightarrow$ **more reactive**

$\Downarrow$

Weaker metallic bonding $\Rightarrow$ **MP decreases**

Group VII: Down the group,

valence shell further from nucleus

$\Downarrow$

more difficult to attract one electron

$\Downarrow$

less reactive

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8. Metals

– Good conductor of Heat and Electricity
– Malleable, Ductile [hence, pure metals are usually soft and weak]
– Most of them have high mp, bp and density

- Alloys
  – mixture of a metal with other elements
  – atoms of different sizes disrupts the orderly arrangement of atoms
  – layers are more difficult to slide and thus stronger and harder than pure metals

Extraction method	Reactivity	Metal	Cold water/ Steam	Hydrochloric acid	Metal oxides with Carbon	Metal oxides with hydrogen	Heat on metal carbonates
Electrolysis of metal ores	<u>Most reactive</u> High tendency to lose valence electrons.	**Potassium**	React with cold water	Decreasing reactivity. Less vigorous reaction.	Carbon **cannot** remove oxygen from their oxides	Hydrogen **cannot** remove oxygen from their oxides	Do not decompose when heated
		Sodium					
		Calcium					Increasing ease of decomposition of metal carbonate. Metal carbonate less stable to heat
		Magnesium	React with steam				
	More reactive metal displaces the less reactive metal from its ions and oxide	**Aluminium**					
Heat metal oxides with carbon (Cheap coke)		**Zinc**			Carbon **can** remove oxygen from their oxides with **increasing ease**.		
		Iron				Hydrogen **can** remove oxygen from their oxides with **increasing ease**.	
		Tin					
		Lead	NO reaction	Very slow even with warm acid			
		Copper		NO reaction			
		Mercury					
Heat metal oxide	<u>Least reactive</u> Low tendency to lose valence electrons.	**Silver**					
Found as element		**Gold**					

Exceptions

- Aluminium seemed to have **no** reaction with acid / steam
  – due to a protective layer of **aluminium oxide**
- When lead react with HCl(aq) or H_2SO_4(aq), reaction stops as soon as it begins
  – a layer of **impervious** (**insoluble)** $PbCl_2$ or $PbSO_4$ is formed

- Extraction of Iron
  – from haematite, Fe_2O_3
  – with coke, limestone, blast of hot air
  – the extracted iron is used to make steel

1. $C\,(s) + O_2\,(g) \rightarrow CO_2\,(g) + \textit{Heat}$
2. $CO_2\,(g) + C\,(s) \rightarrow 2CO\,(g)$
3. a) $Fe_2O_3\,(s) + 3CO\,(g) \rightarrow \mathbf{2Fe\,(\mathit{l})} + 3CO_2\,(g)$
   b) $Fe_2O_3\,(s) + 3C\,(s) \rightarrow 2Fe\,(l) + 3CO\,(g)$

Extract iron (1–3)

4. $CaCO_3\,(s) \rightarrow CaO\,(s) + CO_2\,(g)$
5. $SiO_2\,(s) + CaO\,(s) \rightarrow \mathbf{CaSiO_3\,(\mathit{l})}$

Remove impuritie (4–5)

- Steel
  – alloy of iron and other elements
  – mild steel, hard steel, stainless steel
- Rusting of Iron / Steel
  – Rust: hydrated iron (III) oxide
  – Require presence of **both** water and oxygen
  – Prevention of rusting by **surface** protection or **sacrificial** protection

9. Acids, Bases and Salts

Acids	Alkalis
• produces H^+ ions when dissolved in water • Sour taste • **Turn blue litmus red** • pH < 7 • Conducts electricity (electrolytes)	• produces OH^- ions when dissolved in water • Bitter taste and slippery feel • **Turn red litmus blue** • pH > 7 • Conducts electricity (electrolytes)
Chemical Properties / Reactions	
1. Acid + Reactive Metals → Salt + Hydrogen gas	1. Alkalis + Acids → Salt + Water
2. Acids + Carbonates → Salt + Water + Carbon dioxide gas	2. Alkalis + Ammonium compounds → Salt + Water + Ammonia gas
3. Acids + Metal oxides → Salt + Water OR, Acids + Metal hydroxides → Salt + Water	3. Reaction with solutions of metal ions to form a precipitate. * *<u>This reaction is used as a test to identify metal ions</u>.*
4. Acids + Alkalis → Salt + Water *H^+ (aq) + OH^- (aq) → H_2O (l)*	

- **Strength** of acid / base depends on
  – **extent of ionisation**/dissociation when dissolved in water

HNO_3 (aq) → H^+ (aq) + NO_3^- (aq)
CH_3COOH (aq) ⇌ H^+ (aq) + CH_3COO^- (aq)
$NaOH$ (aq) → Na^+ (aq) + OH^- (aq)
NH_3 (g) + H_2O (*l*) ⇌ NH_4^+ (aq) + OH^- (aq)

Base
Soluble in water?
Yes → **Alkali** → ***NaOH, KOH, aq. NH_3***
No → **Insoluble Base** → ***e.g. CuO, ZnO, etc.***

- Concentration depends on the amount of solute dissolved in a solution

Strength ≠ Concentration

Red	Orange	Yellow	Green	Blue	Indigo	Violet
pH: 0 ←			**7**			→ **14**
Most acidic			Neutral			Most alkaline
Strong acids			Pure water			Strong alkalis
High H^+ ion concentration						Low H^+ ion concentration
Low OH^- ion concentration						High OH^- ion concentration

Oxides

Type of oxide	Examples	Reactions
Basic	Metallic oxides [except ZnO, Al_2O_3, PbO]	React with acids to form salt and water
Amphoteric	**Z**nO, **Al**$_2$O$_3$, **P**bO	React with both acid and alkali to form salt and water
Neutral	H_2O, NO, CO	No reaction with acid and alkali
Acidic	Non-metal oxides [except H_2O, NO, CO]	React with alkalis to form salt and water

Solubility table

Type of compounds	Insoluble	Soluble
Oxides	All except	Group I and Ammonium
Hydroxides	All except	Group I and Ammonium *[$Ca(OH)_2$ sparingly soluble]*
Carbonates	All except	Group I metal and Ammonium
Sulfates	Calcium, Barium, Lead	The rest
Chlorides (Halides)	Silver, Lead	The rest
Nitrates	None	All
Group I and Ammonium	None	All

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Preparation of Salts

	Titration	React acid with EXCESS insoluble substances	Precipitation
Solubility of salt	Soluble	Soluble	Insoluble
Type of salt	Group I and Ammonium salts	Soluble salts [except Group I and Ammonium salts]	All insoluble salts
Solubility of starting material	Two aqueous solutions	One aqueous solution and One insoluble substance in excess	Two aqueous solutions
Starting material	Acid + Alkali Acid + Soluble carbonates	Acid + Reactive metals Acid + Insoluble metal oxides/hydroxides Acid + Insoluble carbonates	One solution containing the required cation and One solution containing the required anion
Examples	*To prepare NaCl* HCl **(aq)** + $NaOH$ **(aq)** $\rightarrow$ $NaCl$ **(aq)** + H_2O (l)	*To prepare $Cu(NO_3)_2$* $2HNO_3$ **(aq)** + CuO **(s)** $\rightarrow$ $Cu(NO_3)_2$ **(aq)** + H_2O (l)	*To prepare $PbSO_4$* **Pb**$(NO_3)_2$ **(aq)** + Na_2**SO_4** **(aq)** $\rightarrow$ **$PbSO_4$ (s)** + $2NaNO_3$ (aq)
Steps	1) Pipette alkali into flask 2) Add indicator 3) Pour acid into burette 4) Add acid into flask until indicator change colour 5) Record volume of acid and alkali added 6) Repeat the titration without indicator 7) Carry out crystallisation process on the resulting salt solution to obtain pure dry crystals	1) Add CuO into acid until some of it cannot dissolve 2) Filter the mixture to remove excess CuO 3) Collect the filtrate 4) Carry out crystallisation process on the resulting salt solution to obtain pure dry crystals	1) Mix the two aqueous solution 2) Filter the mixture 3) Collect the residue 4) Wash with distilled water 5) Dry the salt between filter papers
Things to take note	Titration is repeated without indicator to ensure that it does not become impurities in the salt after crystallisation.	Excess insoluble substance must be added to ensure that all acid has reacted so that acid will not contaminate the salt after crystallisation.	

Identification of Cations

Cations	Aqueous NaOH	Excess Aqueous NaOH	Aqueous Ammonia	Excess Aqueous Ammonia
Cu^{2+}	Blue ppt [$Cu(OH)_2$]	ppt insoluble	Blue ppt [$Cu(OH)_2$]	ppt **soluble** and form deep blue solution
Fe^{2+}	Green ppt [$Fe(OH)_2$]	ppt insoluble	Green ppt [$Fe(OH)_2$]	ppt insoluble
Fe^{3+}	Red-brown ppt [$Fe(OH)_3$]	ppt insoluble	Red-brown ppt [$Fe(OH)_3$]	ppt insoluble
Ca^{2+}	White ppt [$Ca(OH)_2$]	ppt insoluble	No ppt.	-
Zn^{2+}	White ppt [$Zn(OH)_2$]	ppt **soluble** *(reaction with NaOH)*	White ppt [$Zn(OH)_2$]	ppt **soluble**
Al^{3+}	White ppt [$Al(OH)_3$]	ppt **soluble** *(reaction with NaOH)*	White ppt [$Al(OH)_3$]	ppt insoluble
Pb^{2+}	White ppt [$Pb(OH)_2$]	ppt **soluble** *(reaction with NaOH)*	White ppt [$Pb(OH)_2$]	ppt insoluble
NH_4^+	No ppt. Ammonia gas formed when warmed		No ppt.	-

*Add magnesium/zinc to test for **H^+ ions**.

Identification of Anions

Anions	Test	Observations
Nitrate [NO_3^-]	Add aq. NaOH, Al powder, warm	Effervescence. NH_3 evolved turns moist red litmus blue
Carbonate [CO_3^{2-}]	Add dilute HNO_3 or HCl	Effervescence. CO_2 evolved forms white ppt with limewater
Sulfate [SO_4^{2-}]	Add dilute HNO_3 and aq. $Ba(NO_3)_2$	White ppt [$BaSO_4$] formed
Chloride [Cl^-]	Add dilute HNO_3 and aq. $AgNO_3$	White ppt [AgCl] formed
Iodide [I^-]	Add dilute HNO_3 and aq. $AgNO_3$	Yellow ppt [AgI] formed

Identification of Gases

Gases	Test	Observations
NO_2	Moist blue litmus paper	Brown, irritating gas, turn moist blue litmus paper red
Cl_2	Moist blue litmus paper	Yellowish green, irritating gas, turn blue litmus paper red and bleached
NH_3	Moist red litmus paper	Colourless, pungent gas, turn moist red litmus paper blue
SO_2	Acidified $K_2Cr_2O_7$	Colourless, irritating gas, change acidified $K_2Cr_2O_7$ from orange to green
CO_2	Bubbled into limewater	Colourless, odourless gas, form white ppt [$CaCO_3$] with limewater
H_2	Lighted splint	Colourless, odourless gas, extinguish lighted splint with pop sound
O_2	Glowing splint	Colourless, odourless gas, rekindles glowing splint

Test for Water

Substance	Test	Observations
Water	Blue cobalt (II) chloride paper.	Blue paper turns pink.

10. Redox Reactions

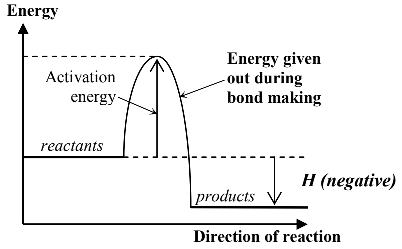
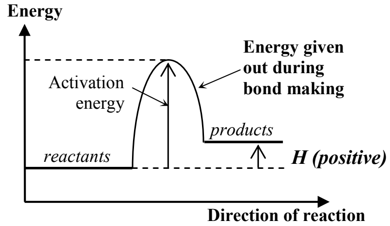
Oxidation	Reduction
▪ **Gain** oxygen ▪ **Increase** in oxidation state ▪ **Loss** hydrogen ▪ **Loss** electron	▪ **Loss** oxygen ▪ **Decrease** in oxidation state ▪ **Gain** hydrogen ▪ **Gain** electron

Calculating oxidation states of elements / elements in a compound

Substance	Oxidation State	Substance	Oxidation State
Free elements	0	Hydrogen in a compound	+1, [except metal hydrides]
Sum of oxidation states of all elements in a compound	0	Oxygen in a compound	-2, [except H_2O_2]
Simple ions	Charge on the ion	Group I elements in a compound	All +1
Sum of oxidation states of all elements in a polyatomic ion	Charge on the ion	Group II elements in a compound	All +2

Oxidising Agent	Reducing Agent
▪ Oxidises other substance ▪ Itself is reduced ▪ Used to test for reducing agents Example: Acidified aqueous potassium dichromate (VI), Chlorine, Acidified aqueous potassium manganate (VII)	▪ Reduces other substances ▪ Itself is oxidised ▪ Used to test for oxidising agents Example: Aqueous potassium iodide, Reactive metals

11. Energy from chemicals

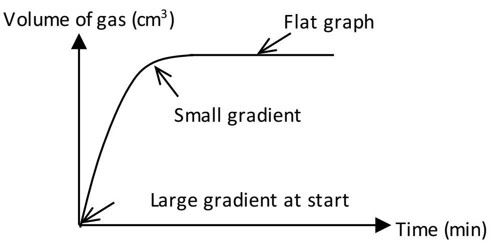
Exothermic Reactions	Endothermic Reactions
• Heat energy is given out • Surrounding temperature increases	• Heat energy is absorbed • Surrounding temperature decreases
• Energy of products < Energy of reactants	• Energy of products > Energy of reactants
• ΔH is negative	• ΔH is positive
• Heat given out in bond making > heat absorbed in bond breaking	• Heat given out in bond making < heat absorbed in bond breaking
• After a heat source is applied, the reaction can continue on its own	• The reaction cannot continue on its own. • Heat energy must be continually applied for the reaction to continue.
 Fig. 3: Energy level diagram for exothermic reaction	 Fig. 4: Energy level diagram for endothermic reaction

- Activation energy
  - the minimum energy needed to start a reaction
  - to break the bonds in the reactants before making new bonds
  - a 'barrier' which must be overcome before any reaction can start

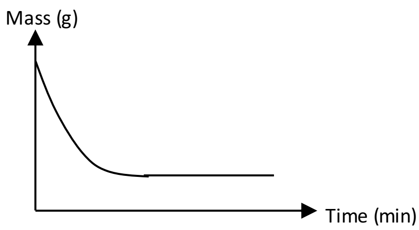
Overall heat change of reaction, ΔH = (Total heat involved in bond breaking) + (Total heat involved in bond making)

12. Speed of Reactions

- Measuring speed of reactions by
  - measure change in volume of gas evolved over time
  - measure change in pressure of gas evolved over time
  - measure change in mass of reaction mixture over time



Volume of gas produced against time



Change in mass against time

- Factors affecting speed of reactions
  - Concentration of solution
  - Pressure of gas
  - Particle size of solid
  - Temperature
  - Presence of catalyst

- The factors increase the speed of reaction by
  - Increasing frequency of collisions and/or
  - Increasing kinetic energy of the reacting particles

13. Atmosphere

Pollutant	Source	Harmful Effect	Solution
Sulfur dioxide	▪ Burning sulfur-containing fuels in power stations and industries $S\ (s) + O_2\ (g) \rightarrow SO_2\ (g)$ ▪ Volcanic eruption	♥ Irritates eye ♥ Breathing problems 🌎 Enters leaves and affect plant growth 🌎 Main cause of acid rain	**Flue gas desulfurisation in power stations** ▪ Powdered limestone is added to the hot gases and decomposes $CaCO_3\ (s) \rightarrow CaO\ (s) + CO_2\ (g)$ $CaO\ (s) + SO_2\ (g) \rightarrow CaSO_3\ (s)$ $2CaSO_3\ (s) + O_2\ (g) \rightarrow 2CaSO_4\ (g)$
Nitrogen oxides (NO_x)	▪ Lightning, Forest fires, Internal combustion engines ▪ Nitrogen and oxygen in air combines at high temperatures $N_2\ (g) + O_2\ (g) \rightarrow 2NO\ (g)$ $2NO\ (g) + O_2\ (g) \rightarrow 2NO_2\ (g)$	♥ Irritate and damage lungs 🌎 Forms photochemical smog 🌎 A cause of acid rain	▪ Fit catalytic converters to cars ▪ In the first half of converter: NOx react with CO as they pass through a (platinum) catalyst $2NO\ (g) + 2CO\ (g) \rightarrow 2CO_2\ (g) + N_2\ (g)$ ▪ In the second half: Air enters and oxidises unburnt hydrocarbon and CO to form CO_2 and water
Carbon monoxide [colourless, odourless]	▪ Forest fires ▪ Incomplete combustion of carbon containing fuels in motor engines, e.g. cars	♥ Reduces ability of haemoglobin in blood to carry oxygen ♥ Causes breathing problems, headache and even death 🌎 Forms photochemical smog	

Acid rain (pH <5) – any rainfall that has an acidity level beyond the acidity of non-polluted rainfall (pH 5.6)
– caused by sulfur dioxide and nitrogen dioxide

$2SO_2\ (g) + O_2\ (g) + H_2O\ (g) \rightarrow H_2SO_4\ (l)$

$4NO_2\ (g) + O_2\ (g) + 2H_2O\ (g) \rightarrow 4HNO_3\ (l)$

Effects of acid rain	Reducing effects of acid rain
▪ Make soil acidic ⇒ many plants cannot grow well ▪ Corrodes buildings and objects made of limestone ▪ Attacks metal ⇒ galvanised iron corrode faster ▪ Damages trees by leaching important nutrients from soil ▪ Acidic water kills fish and other aquatic life	▪ Burn fuels containing little or no sulfur. E.g. use natural gas ▪ Use catalytic converters in cars ▪ Remove acidic gases in power stations ▪ Neutralise the acids in lakes with calcium carbonate and acidic soil with slaked lime (CaOH)

Carbon cycle – involves the processes: combustion, respiration and photosynthesis

Increase emission of carbon dioxide and methane gases increases greenhouse effect and global warming

Effects of Global Warming	Preventing effects of Global Warming
▪ Seawater expands and sea levels rise ▪ Polar ice caps melt and sea levels can rise by over 20 m ▪ Big changes in global climate ⇒ Floods and droughts, serious shortage of food	▪ Reduce usage of fossil fuels ▪ Use more clean methods of producing electricity ▪ Use of electric vehicles ▪ International agreements e.g. the Kyoto Protocol in 1997