

Jack's Chemistry Revision Notes

or...

Why you'd have to be mad to study chemistry

VERSION 0.2.0 – 16 MAY 1999

Atomic Structure (1.1)

A – mass number

Z – number of protons

Mass spectrometer [1]

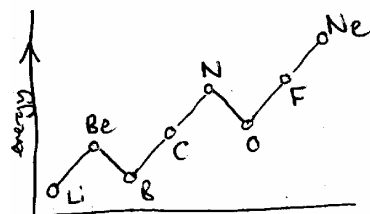
Vaporised sample is **ionised**, **accelerated** by an electric field, **deflected** by a magnet and made to hit a **detector**.

Order of electron shells [2]

1s 2s 2p 3s 3p 4s 3d 4p

[Ar] = $1s^2 2s^2 2p^6 3s^2 3p^6$

Ionisation energy across period 2 [3]



1. Shielding by existing electron orbitals causes this.

eg. Boron shielded by whole 2s orbital.

2. Pair effects

Nitrogen has 3 unpaired p electrons.

Oxygen has 2 unpaired p electrons, and a pair. It is easier to remove one of the p electrons from oxygen than from nitrogen, because there are **less unpaired electrons**.

- Ionisation energy **decreases down** a group. (Shielding effect of previous shells)
- Ionisation energy generally **increases from left to right** across a group, as nucleus becomes more positive but electrons are still at the same distance (greater attraction).

Amount of Substance (1.2)

Misc [4]

Avogadro Constant = 6.0×10^{23} per mole

A_r = relative atomic mass

M_r = relative molecular mass

M_r is measured relative to Carbon 12.

Ideal Gas Equation [5]

$$pV = nRT$$

($R = 8.31 \text{ JK}^{-1} \text{ mol}^{-1}$)

Bonding (1.3)

Bond types [6]

Co-ordinate: A covalent bond where both electrons are from the same atom.

Metallic: lattice of +ve ions surrounded by delocalised electrons.

Pauling's scale of electronegativity [7]

Scale from 0.7 to 4.0.

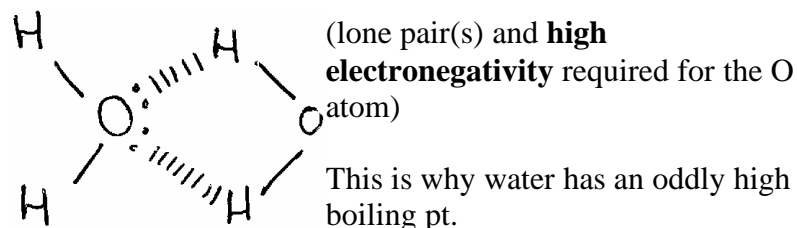
electronegativity is the power of an atom to withdraw e^- density from a covalent bond

Covalent bonds are not always symmetrical: they may be polarised:

- **NEGATIVE anions** can be polarised by **cations** (+ve) of high charge density.
eg. HCl is polar. The H is more positive ($\delta+$) than the Cl.
- Ionic bonds have covalent character unless the ions are perfectly spherical. The greater the charge to size ratio of the ions is, the more polarised the ionic bond is, and the more covalent character it will show.

eg. **NaCl** is truly ionic, **MgCl₂** has some covalent character, and **AlCl₃** and **SiCl₄** are covalent.

Intermolecular bonds: Hydrogen bonding [8]



($\text{H}_2\text{S} = -50^\circ\text{C}$, $\text{H}_2\text{Se} = -30^\circ\text{C}$, but $\text{H}_2\text{O} = 100^\circ\text{C}$)

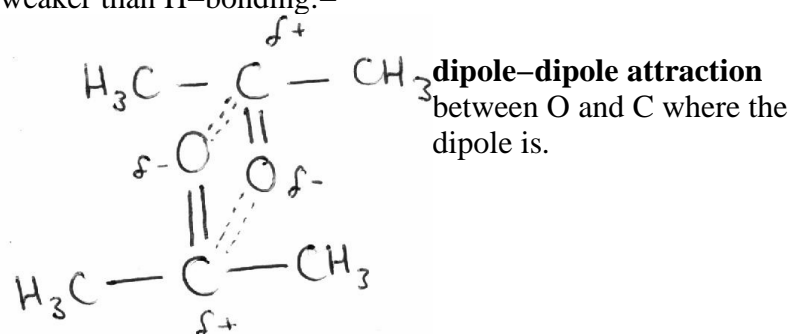
H-bonds are **very strong** (less than covalent bonds though) and are called **permanent** bonds.

NH_3 can make 1 hydrogen bond, HF can make 3.

How it works: electrons on the H are pulled away by the high electronegativity of the other atom, leaving an exposed side of the proton, which attracts negative areas on other molecules, eg. Lone pairs.

Intermolecular bonds: Dipole-dipole attraction [9]

weaker than H-bonding:-



Intermolecular bonds: Van der Waals forces [10]

Non-polar molecules do have I/M forces, they are the **induced dipole-dipole** type.

e^- movement induces a dipole in a molecule. These forces are attracting and breaking all the time.

Bond Rules

All things have **Van der Waals** I/M forces. If a molecule has a **difference in electronegativity** across it (i.e. **A dipole**) then there will also be **dipole-dipole attraction**. And if one of the atoms has **very high electronegativity** and **lone pairs** (N, O, F), and is **bonded to a hydrogen atom**, then **hydrogen bonding** will also exist.

Types of crystal [11]

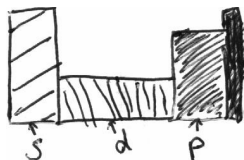
- **Ionic** crystals (like NaCl) have a large cubic structure. (and a high melting point)
- **Molecular** (giant covalent) eg. Graphite, diamond
Every atom shares electrons with neighbouring atoms.
- **Giant Atomic** – just atoms packed as tightly as possible – stacking like balls in pyramid structure. This is either **cubic close packing** or **hexagonal close packing**.
(lower melt point than ionic crystals)

Repulsion rule [12]

A **lone pair of electrons** produces a **greater repulsion** effect than a **bonding pair**.

lone pair – lone pair repulsion > **lone pair – bonding pair** repulsion > **bonding pair – bonding pair** repulsion

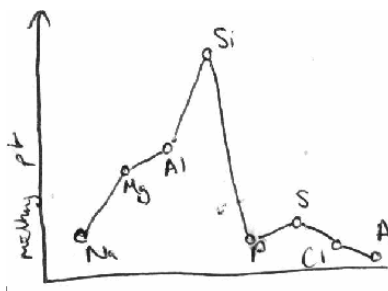
Periodicity (1.5)



Patterns in period 3 [17]

Period #3 is a bit like period #2.

- Atomic radius **decreases** from left to right (**Na larger than Ar**)
(increased nuclear charge => increased nuclear size => pull on e^- is greater)
- melting pt. increases from Na to Si, then drops for P, increases slightly for S, then decreases again for Ar:

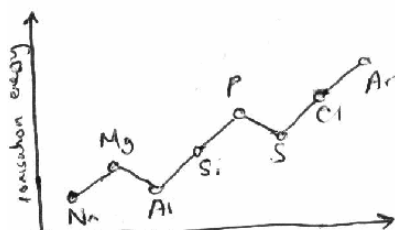


– For metals, melting point increases as the charge/size ratio increases – better metallic bonding.

– Silicon is **macromolecular**, so it has very strong bonds. (covalent bonds linking all atoms together)

– The final elements are simple molecular: with weak Van der Waals I/M forces.

- **electronegativity** increases from left to right.
(increasing nuclear charge => electrons attracted more strongly)



- **first ionisation energy** increases from left to right (same nuclear size, but greater nuclear charge, so electrons harder to remove)

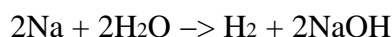
Extra Solid Types

PCl₅ actually exists as **an ionic solid** consisting of PCl₄⁺ and PCl₆⁻. **AlCl₃** is a solid **dimer** at room temperature.

Patterns down group 1 [18]

- radius increases (more e⁻ shells)
- ionisation energy decreases (more e⁻ shells)
- electronegativity decreases (greater shielding effect)
- **melting point decreases** (metal-metal bond strength decreases)

Group I metals react with **water** to form **metal hydroxides** and **hydrogen**:



Metal oxides and chlorides [19]

MgO, SO₂, MgCl₂, and AlCl₃ can be formed by direct combination (burning).

Reactions of oxides of period 3 elements with water [20]

Basic element	Na ₂ O	Dissolves slowly to form strong alkali
	MgO	Dissolves slowly to form weak alkali
Amphoteric element	Al ₂ O ₃	Does not react
Acidic element	SiO ₂	Does not react
	P ₄ O ₁₀	Forms phosphoric acid soln.
	SO ₂	Forms sulphuric acid soln.
	SO ₃	Forms sulphuric acid soln.

So, the **metals** form alkalis, and the **non-metals/metalloids** either form acids or do not react.

Reactions of chlorides of period 3 elements with water [21]

NaCl	Dissolve easily, solution pretty neutral
MgCl ₂	
Al ₂ Cl ₆	Reacts vigorously with water, produces HCl fumes
SiCl ₄	
PCl ₅	Solution acidic , pH 2 or 3.

Additional Notes on periodicity [22]

In each period the oxides of the **metals** and **metalloids** have **giant structures**, whereas the oxides of **non–metals** are composed of **simple molecules**.

- **The simple molecules** (non–metals) react **much more readily** than giant structures.

From **left** to **right**, the oxides of all these molecules change from being **ionic, involatile, and metallic**, through being **giant molecular, involatile and amphoteric**, to being **simple molecular, volatile, and acidic**.

Lithium (thermal stability of compounds) [23]

- Lithium Carbonate is **much less stable** than other group 1 carbonates.
- Lithium Hydroxide is also **less stable**.
- LiNO_3 decomposes on heating to $\text{LiO}_2(\text{s})$ but all other group 1 metals decompose to their corresponding **nitrates**, eg. $\text{NaNO}_2(\text{s})$.
- Lithium compounds have **more covalent character** than other Group 1 compounds, eg. **LiI** is soluble in methanol and propanol.

The Halogens (2.4)

[40]

Most electronegative: **Fluorine**
 Highest boiling point: **Astatine**
 Best oxidising agent: **Fluorine**
 Best reducing agent: **Astatide ions**
 (I/M forces are just Van der Waals)

Products of reactions between NaX and Sulphuric Acid [41]

(where X is a halogen): $\text{NaX} + \text{H}_2\text{SO}_4(\text{conc})$

Halide ions can reduce sulphur in H_2SO_4 by varying degrees. In H_2SO_4 the S has oxidation state +6. Iodide ions are able to reduce this to +4 (SO_2), 0 (S) and -2 (H_2S).

reagent	product
NaF	HF
NaCl	HCl
NaBr	HBr and a little Br_2 <i>and</i> SO_2

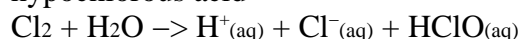
reagent	product
NaI	I ₂ and a little HI <i>and S, H₂S, SO₂</i>

NaBr and NaI are oxidised in the reaction, but HF and HCl cannot be oxidised by the acid.

Br₂ is a brown gas. I₂ is a purple gas, and a black precipitate.

Reactions of Cl

[42] ..with water: chlorine + water → hydrochloric acid + hypochlorous acid



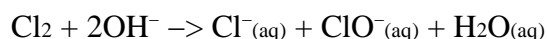
Cl is simultaneously reduced and oxidised – a disproportionation.

The product is also known as chlorine water. It is used to make bleach and treat water, and kills germs.

[43] ..with NaOH

This reaction depends on concentration and temperature.

At **room temperature** and **using dilute NaOH:**



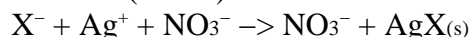
Forms **Sodium Hypochlorite** which is a bleach.

Testing for halides [44]

Silver nitrate (AgNO₃) can be used to detect halide ions.

halide ion + silver nitrate → nitrate ion + silver halide

or, for reaction with X⁻ (halide)



X can be identified from the colour of the silver halide AgX:

[45]

AgI_(s) is **yellow**

AgBr_(s) is **creamy colour**

AgCl_(s) is **white**

X can also be identified by dissolving the silver halide in ammonium hydroxide: [46]

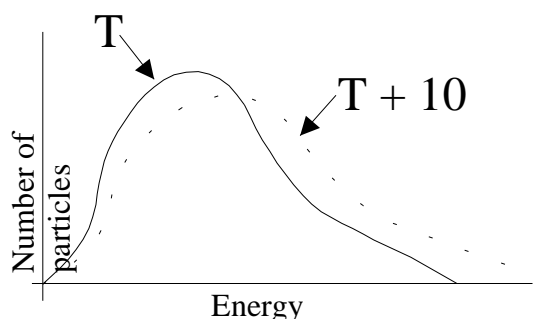
AgI_(s) **doesn't dissolve**

AgBr_(s) **partially dissolves**

AgCl_(s) **dissolves fully**

Kinetics (3.1)

Maxwell–Boltzmann Distribution [47]



increasing T shifts to right

Area under curve is total number of particles

Reaction rate [48]

Rate = Δ amount / Δ time or Δ concentration / Δ time

The rate is affected by:

1. **state of division** (i.e. The surface area, powdered form is faster)
2. **temperature** (more energy \rightarrow more collisions have required E_a)
(*a small increase in T may lead to a large increase in rate*)
3. **concentration** (increased P(collisions))
4. **catalyst** (provides alternate route with lower E_a)

Rate equation [50]

Rate = $k [A]^m [B]^n$

m and **n** are the **orders** of the reaction with respect to reagents A and B.

k is the **rate constant**.

Thermodynamics & Energetics (4.1 1.4)

Enthalpy [74]

Definition: Enthalpy (**H**) is the energy content **at constant pressure**.

Standard enthalpy changes (ΔH^\ominus) refer to standard conditions: 1

atm, 298K, 1M

Types of Enthalpy [75]

+ve means that energy goes from surroundings to system (i.e. Endothermic)

Ionisation enthalpy	One mole of e⁻ from gaseous atoms	ΔH^{\ominus}_{I1}	+ve
Electron affinity	One mole of e⁻ added to gaseous atoms	ΔH^{\ominus}_{e1}	-ve
Lattice dissociation enthalpy	Ionic solid dissociating to gaseous atoms	ΔH^{\ominus}_{lat} t	+ve
Solution enthalpy	One mole of ionic solid dissolving in water to make aqueous ions	ΔH^{\ominus}_{sol} n	+ve
Enthalpy of hydration	1 mole Gaseous ions forming aqueous ions	ΔH^{\ominus}_{hy} d	-ve
Enthalpy of combustion	Buring 1 mole of reagents in standard states	ΔH^{\ominus}_c	-ve
Enthalpy of formation	Forming 1 mole of something in standard state from elements in standard states	ΔH^{\ominus}_f	
Enthalpy of sublimation	Making 1 mole of gaseous atoms	ΔH^{\ominus}_{su} b	+ve
Enthalpy of atomisation	Often the same as sublimation.	ΔH^{\ominus}_{dis} s	+ve

Hess's Law [14]

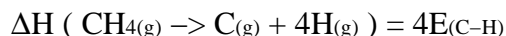
Total Δ enthalpy is independent of the reaction route taken.

Bond Enthalpy [76]

Bond Enthalpy is the definite amount of energy associated with each chemical bond. Bond enthalpies can be used to **predict** whether or not, and how easily two substances will react.

bond breaking – endothermic bond making – exothermic

example: all the bonds in CH₄ are identical C–H bonds with the same bond enthalpy, E_(C–H). To break all four, 4E_(C–H) is required.



Generally, the lower the bond enthalpy, the weaker the bond.

- **But E_{bond} is not constant, it varies according to the environment of the bond. The same bond in different compounds will have different values of E_{bond} .**

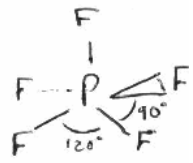

When **heat** or **light** can break bonds, it will usually break the **weakest** (lower bond enthalpy) first.

eg. Cracking works because C–C bonds are weaker than H–C bonds (C–C bonds split homolytically to form radicals).

eg. $\text{H}_2 + \text{Cl}_2$ works because Cl–Cl bonds split in the presence of light. (initiation, propagation, termination steps..)

Molecule shapes [16]

$\text{H}-\text{C}\equiv\text{C}-\text{H}$	Linear	180
$\text{H}-\ddot{\text{O}}-\text{H}$	V-shaped	105
CH_4	Tetrahedral	109.5
$\text{H}-\ddot{\text{N}}-\text{H}$ H	Pyramidal	107
$\text{F}-\text{B}-\text{F}$ F	Planar	120
$\text{H}-\ddot{\text{N}}-\text{H} \rightarrow \text{B}-\text{F}$ H	Both tetrahedral	109.5

	Trigonal bipyramidal	90 and 120
	Octahedral	90

Spontaneous reactions [78]

ΔH , whilst important, is not sufficient to explain spontaneous change.

=> spontaneous exothermic change (eg. burning) makes sense in terms of ΔH but spontaneous endothermic change does not.

Entropy (S) [79]

UNITS: $\text{JK}^{-1}\text{mol}^{-1}$ (NOT kJ)

Positive ΔS for a reaction indicates that the reaction is **feasible** (ie. spontaneous).

S rules of thumb [80]

- S of gases > S of liquids > S of solids
- ions and molecules in solution have higher S than solids
- larger molecules breaking down into smaller molecules show an increase in entropy
- increasing T increases S
- A change of state has much more effect on S than a change of temperature
- The more complex a molecule is, the higher S is (at same temperature) because there are more bonds.

2nd law of thermodynamics [81]

$$\Delta S_{\text{total}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$$

ΔS_{total} **must** be **positive** because **overall entropy is always increasing**.

And because $\Delta S_{\text{system}} = \Delta \Sigma S_{\text{products}} - \Delta \Sigma S_{\text{reactants}}$
and $\Delta S_{\text{surroundings}} = -\Delta H / T$

the feasibility of a reaction can be determined by calculating the ΔS_{system} and $\Delta S_{\text{surroundings}}$ from databook values. If ΔS_{total} is

positive, the reaction is feasible at that temperature. Thus **the feasibility of a reaction depends on the balance between entropy and enthalpy.** [82]

Gibbs free energy change [83]

$$\Delta G = \Delta H_{\text{system}} - T\Delta S_{\text{system}}$$

For a feasible reaction ΔG **must be zero or negative**, note **this is the opposite way round to ΔS .**

Equilibria (4.2 2.1)

Many reactions are reversible. At equilibrium, both forward and reverse reactions are proceeding. For a homogenous system in equilibrium, where: $A + B \rightleftharpoons C + D$,

$$K_c = \text{Right} / \text{Left} = [C]_{\text{eq}}[D]_{\text{eq}} / [A]_{\text{eq}}[B]_{\text{eq}}$$

[25] Le Chatelier's principle: The system resists change.
=> Increase conc. of something, it reduces it.

- **Catalyst has no effect on K_c .**
- **Changes in overall pressure have no effect on K_c .**

But K_c depends on temperature. The effect of T can be clearly defined:

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

- where R = gas constant (8.31)

Partial pressure and K_p [85]

=> partial pressure is analogous to concentration

The partial pressure of a gas in a gas mixture is the pressure that would be exerted if that gas alone filled the whole volume occupied by the mixture (Dalton's law)

partial pressure $P_A = P_{\text{total}} \times \text{mole fraction}$
(P_A of oxygen in air is 0.2 atm, as this is 1/5 of 1 atm (4 parts nitrogen, 1 part oxygen))

K_p is the equilibrium constant for a gas system:

$$K_p = \text{Right} / \text{Left} = P_{(C)\text{eq}}P_{(D)\text{eq}} / P_{(A)\text{eq}}P_{(B)\text{eq}}$$

Note that changing **pressure** of the system does **not** affect K_p , but changing **temperature** does affect K_p .

Contact Process [26]

- Makes SO_3 from SO_2 by adding O_2 .
Low temperature, high pressure, platinum/manganese catalyst
(typically 450°C , and a few atm – not really low T/high p, but obtains high yield at low expense)

Haber Process [27]

- makes ammonia from N_2 and H_2 . Low T and high P.
(typically 500°C , and **100–300** a.t.m., with **iron catalyst**)
The iron catalyst is supported on **silica**.

Acid–Base Equilibria (4.3 2.2)

Bronsted–Lowry: acid is a proton donor, base is a proton acceptor

In any acid–base reaction a proton is transferred.

pH [30]

$$\text{pH} = -\log [\text{H}^+] \quad (\text{H}^+ \text{ in mol dm}^{-3})$$

Dissociation of water [31]

Water is very weakly dissociated: $K_c = \frac{[\text{H}^+][\text{OH}^-]}{[\text{H}_2\text{O}]}$

The ionic product of water, $K_w = [\text{H}^+][\text{OH}^-]$

At 25°C **only**, $K_w = 10^{-14} \text{ mol}^2\text{dm}^{-3}$.

pH of a strong base found by finding $[\text{H}^+]$: $K_w = [\text{H}^+] \times [\text{OH}^-]$.

- weak acids dissociate only partially in aq. soln.
- strong acids/bases are fully dissociated: and fully ionised in solution

>>> Dissociations are always written with ions on the right.

Therefore, as $K_c = \text{right} / \text{left}$, K_a , K_c etc are always **ions / molecules**.

– Ions Right –

Monoprotic weak acids

A weak acid is only partially dissociated: $\text{HA} \rightleftharpoons \text{H}^+(\text{aq}) + \text{A}^-(\text{aq})$

The dissociation constant, K_a , is found by:

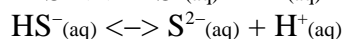
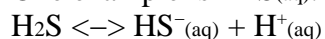
$$K_a = \frac{[\text{H}^+(\text{aq})][\text{A}^-(\text{aq})]}{[\text{HA}]}$$

The higher the value of K_a , the stronger the acid.

Diprotic weak acids

These are like monoprotic weak acids, except they dissociate twice, producing 2H^+ .

One example is $\text{H}_2\text{S}(\text{aq})$:



There are two values of K_a , K_{a1} , and K_{a2} :-

$$K_{a1} = \frac{[\text{HS}^-][\text{H}^+]}{[\text{H}_2\text{S}]}$$

$$K_{a2} = \frac{[\text{S}^{2-}][\text{H}^+]}{[\text{HS}^-]}$$

K_{a2} is always much less than K_{a1} .

Finding K_a

K_a may be found by this approximation, as $[\text{H}^+] \approx [\text{HA}^-]$, and $[\text{H}^+] \approx 0$. (as very few H^+ come from the water that the acid is in)

$$K_a \approx \frac{[\text{H}^+]^2}{[\text{HA}]}$$

Finding pH

and so, if $[\text{HA}]$ and K_a are known for a weak acid, it's pH can be calculated:

$$[\text{H}^+] \approx \sqrt{([\text{HA}] \times K_a)}$$

$$\Rightarrow \text{pH} \approx -\log_{10} \sqrt{([\text{HA}] \times K_a)}$$

pKa

$\text{p}K_a$ is a pH-like scale for K_a .

$$\text{p}K_a = -\log_{10} K_a$$

pH and pKa

You can combine the above two equations to make:

$$\text{pH} \approx \frac{1}{2}\text{p}K_a - \frac{1}{2}\log [\text{HA}]$$

Weak bases

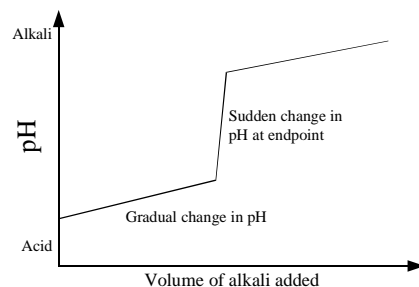
Just as $K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$ for weak acids,

$K_b = \frac{[\text{B}^+][\text{OH}^-]}{[\text{BOH}]}$ for weak bases.

$$K_b \approx \frac{[\text{OH}^-]^2}{[\text{BOH}]}$$

$$\text{pH} \approx \frac{1}{2}\text{pK}_a + \frac{1}{2}\text{pK}_w + \frac{1}{2}\log [\text{BOH}]$$

pH Curves



All acid/base titrations produce a pH graph like the one on the left. (a mirror image for an acid added to an alkali)

The **endpoint** is where the titration should finish: where equal amounts of acid and base

have reacted.

Also known as the **equivalence point** where acid and base are equivalent.

Indicators change colour within a narrow pH range. The right indicator has to be used to show when the titration reaches the endpoint. This isn't always $\text{pH} = 7$ though.

For a strong acid and strong alkali, endpoint is about $\text{pH} 7$.

For a weak acid and strong alkali, endpoint is greater than $\text{pH} 7$.

For a strong acid and weak alkali, endpoint is less than $\text{pH} 7$.

For a weak acid and weak alkali, endpoint is about $\text{pH} 7$.

The indicator should show when the mixture is close to the endpoint.

indicator	Changes between...
Methyl Orange	3 and 4.6
Phenolphthalein	8 and 10

Buffers

- an **acid buffer** consisting of a weak acid (eg. Ethanoic acid) and a salt of the acid (eg. Sodium ethanoate).
- An **alkaline buffer** consisting of a weak alkali (eg. ammonium hydroxide) and a salt of the alkali (eg. ammonium chloride).

Acid buffer: If H^+ ions are added, they will react with the ethanoate ions from the salt (because there are more of them than ethanoate ions from the partially dissociated acid).

Undissociated ethanoic acid will form. If OH^- ions are added, they will combine with the H^+ ions to form water, and more of the acid will dissociate.

Alkaline buffer: same sort of mechanism

Uses of buffers

- calibration of pH meters
- biology eg. Maintaining correct conditions for enzymes
- medicines
- dyes
- electroplating
- injections.

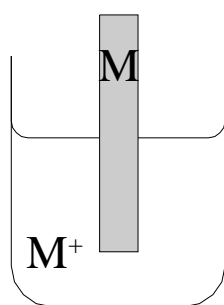
Redox Titration Facts

MnO_4^- reduces to Mn^{2+}

$\text{Cr}_2\text{O}_7^{2-}$ reduces to 2Cr^{3+}

Redox Equilibria (4.4)

The **anode** is where the **oxidation** reaction occurs.



Usually, electrodes are a sample of a metal in a solution of the same metal ions. If $M = \text{Zn}$ and soln is $\text{ZnSO}_4(\text{aq})$:

$\text{Zn}(\text{s}) \mid \text{ZnSO}_4(\text{aq})$

or

$\text{Zn}(\text{s}) \mid \text{Zn}^{2+}(\text{aq})$

both describe this electrode.

Gas electrode

Sometimes a gas electrode must be used. Here the gas under test is bubbled over a platinum electrode in a solution of ions of the same element as the gas. eg. The standard hydrogen electrode is described as:

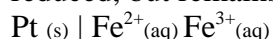
$\text{Pt}(\text{s}) \mid \text{H}_2(\text{g}) \mid \text{H}^+(\text{aq})$

Standard

The standard for half cells is that **everything happens under standard conditions** (1 atm for gases, 298K, and 1M solutions).

Redox electrodes

This is where something in solution is either oxidised or reduced, but remains in solution. eg.



→ a solution containing 1M iron (II) and 1M iron (III) ions. e^- produced by oxidation of Fe^{2+} or lost by reduction of Fe^{3+} enter or leave the solution by the Pt electrode.

Cell reactions – standard layout



- The | is a phase boundary.
- The || indicates a salt bridge.

The oxidation cell is put on the left. The **overall reaction** where oxidation of the left cell occurs is **feasible** if the left cell potential is less than the right cell potential.

Word Association

Oxidation occurs at the **Anode** on the **Left** which is **Negative**

POSITIVE POTENTIALS OXIDISE OTHERS

ie. the **best oxidising agents** have **the highest E values**.

Standard Cell Potential

Standard cell potential, E^\ominus is measured with the Standard Hydrogen Electrode (SHE) as the left hand electrode, and standard conditions. Cells with **high E values** are the best **oxidising agents**: they accept e^- easily.

Calculating E

$$E_{\text{cell}} = E_{\text{right}} - E_{\text{left}}$$

E_{cell} is usually calculated by putting the most positive cell on the right.

Secondary standard

A **calomel** electrode with an E^\ominus value of **+0.27V** is often used in place of the SHE.

POSITIVE POTENTIALS OXIDISE OTHERS

Principles of Catalytic Action (4.5)

Heterogeneous catalytic action

A heterogeneous catalyst is in a different state to the reagents. One (or more) of the reagents is adsorbed onto the catalyst surface, making it more likely to react with the others. For example, it may become **more accessible** to collisions, or be held in a particularly **active configuration**, or be broken into **more reactive fragments**.

The strength of adsorption is very important. **Tungsten adsorbs too well. Silver adsorbs poorly.** So Ni and Pt are more commonly used.

A **support medium** minimises the amount of catalyst while maximising the catalyst surface area.

Homogenous catalytic action

The reaction goes through an intermediate species. If a transition metal catalyst is used, its oxidation state changes.

Specificity

Catalysts are often highly specific. Enzymes will work for only one reaction. Acid/base catalysts (eg. Acid used to catalyse hydrolysis of an ester) are much more general.

Some Catalysts [35]

Transition metals often make good catalysts (change in oxidation state)

Iron → **Haber process**

MnO₂ → decomposition of **H₂O₂**

Ni → margarine (**hydrogenation** of vegetable oil)

Extraction of Metals (5.1)

Extraction of metals usually involves reduction of a metal oxide. There are several reduction methods: the one used depends on the **cost** of the **reductant**, the **energy requirements**, and the **purity of the metal** required.

Reduction of metal oxides with carbon

eg. production of iron.

1. Iron oxide is first **heated** with **coke** (mostly carbon) in a sintering plant.

2. Poured into a blast furnace along with more coke.
3. The carbon reduces the iron oxide to iron:

$$2\text{Fe}_2\text{O}_3 + 3\text{C} \rightarrow 4\text{Fe} + 3\text{CO}_2$$
and also, CO produced by incomplete combustion of coke:

$$\text{Fe}_2\text{O}_3 + 3\text{CO} \rightarrow 2\text{Fe} + 3\text{CO}_2$$
4. The impurities (slag) float on top of the liquid iron poured out of the blast furnace.

Waste gases are used to heat the furnace.

Iron can be further purified by:

- bubbling oxygen through it in a converter to remove carbon
- adding magnesium to remove sulphur
- adding lime (CaCO_3) to remove other impurities (reacts with SiO_2 , the main impurity, producing slag, CaSiO_3)

Carbon reduction is **not perfect for all metals**, as with some others carbides are formed.

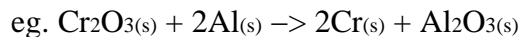
Reduction of metal oxide with an active metal

Mainly used for **chromium** and in **Thermite** process.

Advantages:

- **no carbides** formed
- **lower temperatures** than carbon reduction
- produces **small quantities of very pure metal**

Sort of substitution of one metal for another:



Reduction of metal oxide by electrolysis

Used for **aluminium extraction** because Al_2O_3 is too stable for carbon reduction: the temp would have to be very high. The **bauxite** is dissolved in **molten cryolite** (Na_3AlF_6) to lower its melting point from **2000°C** to **950°C**. Carbon electrodes are used.

Cathode: Al reduced from Al^{3+} to Al

Anode: Oxide ions oxidised to oxygen

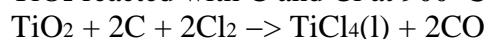
This is a continuous process using much electricity. Cryolite is damaging to the environment.

Reduction of a metal halide by another metal

Used for production of Ti, which is brittle unless very pure.

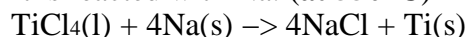
This process produces **very pure metal** in **quantity**. Main Ti ore: TiO_2 .

TiO_2 reacted with C and Cl at **900°C** to form TiCl_4 :



TiCl₄ is a **dangerous product**, it hydrolyses easily and forms HCl fumes in moist air.

It is reacted with Na: (at 550°C)



This is also a **dangerous reaction** as it's highly **exothermic**. (temp rises to 1000°C) The temperatures are so high that the reaction must happen in an inert atmosphere of argon so that titanium oxides are not formed again. It's all very expensive.

- dangerous process
- high temperatures needed
- sodium and chlorine needed
- inert atmosphere used

additional: H₂ can be used as a reducing agent for metals: no carbides are formed.

Transition Metals (5.2 2.3)

General Properties [34]

1. coloured compounds
2. variable oxidation state
3. form complex ions
4. show catalytic activity

A transition metal has an incomplete d sub-shell in at least one of it's oxidation states. The properties arise from this.

The odd order of e.c.'s [37]

- Sc is [Ar] + 4s² + 3d¹

Most of them add 1 normally, eg.

- Ti is [Ar] + 4s² + 3d²

except Chromium and Copper

- Cr is [Ar] + 4s¹ + 3d⁵
- Cu is [Ar] + 4s¹ + 3d¹⁰

On ionisation, the **4s shell goes first, then the 3d shell.**

Complex formation [39]

complex: central metal ion surrounded by ligands

co-ordination number: number of ligands

[Co(H₂O)₆]²⁺ and [Cu(H₂O)₆]²⁺ are **octahedral**.

[CoCl₄]²⁻ and [CuCl₄]²⁻ are **tetrahedral**.

The Colour Table

<u>Initial Ion</u>	<u>Initial Colour</u>	<u>+HCl Colour</u>	<u>+NH₃ Colour</u>
[Cu(H ₂ O) ₆] ²⁺	Blue	Green	Royal Blue
[Co(H ₂ O) ₆] ²⁺	Pink	Blue	
[Fe(H ₂ O) ₆] ²⁺	Green		
[Fe(H ₂ O) ₆] ³⁺	Very Pale Violet		
[Cr(H ₂ O) ₆] ³⁺	Ruby		Purple

The colour of ions

is affected by a **change in ligands**, a **change in co-ordination number**, or **change in oxidation state**.

Ligands

A ligand is an **electron pair donor**.

Unidentate: one tooth: NH₃, Cl⁻

Bidentate: two teeth: ethanedioate ions

Polydentate: many teeth: EDTA, and haemoglobin

Reactions of inorganic compounds in aqueous solutions (5.3)

A **Lewis acid** is an **electron pair acceptor** and a **Lewis base** is an **electron pair donor**.

(A transition metal is one with a partially filled d sub-shell)

Metal ions form metal-aqua ions in water: [M(H₂O)₆]²⁺.

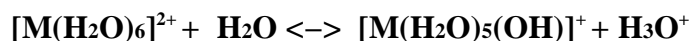
These ions can be present in the solid state: FeSO₄·7H₂O.

Acidity reaction (hydrolysis)

- M³⁺ ions are more acidic than M²⁺ ions (M³⁺ ion has greater polarising power)
- M³⁺ precipitates are amphoteric
- CO₃²⁻ ions react with H₃O⁺ to form carbonic acid. M²⁺ are less acidic than this, and so are unable to displace the CO₂ from it, and that's why they form metal carbonates, whereas M³⁺ are more acidic, and form CO₂ and metal hydroxides.

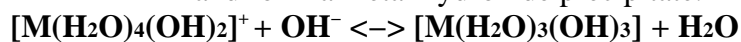
with water...

M²⁺ and M³⁺ ions can lose protons to water to the extent that they have +1 charge:



with OH^- ions...

M^{2+} and M^{3+} ions lose protons as before, but this time can go a step further and form a metal hydroxide precipitate.



M^{3+} pptes can redissolve in OH^- ions to form $[\text{M}(\text{H}_2\text{O})_2(\text{OH})_4]^-$ ions.

M^{2+} pptes do not redissolve.

with CO_3^{2-} ions...

M^{2+} ions form **insoluble metal carbonates** (MCO_3)

M^{3+} ions form **metal hydroxides** and **CO_2 gas**

with NH_3 ions...

M^{2+} and M^{3+} form **hydroxides** and **then substitution occurs.**

Metal Hydroxide Table

<i>Metal ion</i>	<i>Metal aqua solution</i>	<i>Metal hydroxide ppte</i>
Iron (II)	Green	Green
Cobalt (II)	Pink	Blue green
Iron (III)	Pale violet	Brown
Chromium (III)	Ruby	Green
Aluminium (III)		White

Substitution Complex Ion Table

- Metal hydroxides are always initially formed. But with some bases in excess, ligand substitution will occur.

Ion..	With conc HCl	With dil NH_3	With conc NH_3	With v. conc NH_3
$[\text{Cu}(\text{H}_2\text{O})_6]^{2+}$ blue	$[\text{CuCl}_4]^{2-}$ green	$[\text{Cu}(\text{H}_2\text{O})_4(\text{OH})_2]$ blue ppte.	$[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ deep blue	$[\text{Cu}(\text{NH}_3)_6]^{2+}$ royal blue

$[\text{Co}(\text{H}_2\text{O})_6]^{2+}$ pink	$[\text{CoCl}_4]^{2-}$ blue	$[\text{Co}(\text{H}_2\text{O})_4(\text{OH})_2]$ green/blue ppte		$[\text{Co}(\text{NH}_3)_6]^{2+}$ straw colour
$[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$ ruby				$[\text{Cr}(\text{NH}_3)_6]^{3+}$ purple
	-ve ligand substitution	Acidity reaction (NH_3 acts as base)		Complete substitution

The Extra Bastard Hard Evil Table

Oxidation state	7	6	5	4	3	2
Vanadium			Orange	Blue	Green	Violet
Manganese	Purple			Dark brown		Pink
Chromium		Orange			Green	Blue

- Vanadium may be reduced by zinc in dilute acid.
- These are all colours of -ate ions, I think.
- Manganese may be reduced by iron

Silver complexes

Silver ions produce linear complexes.

$[\text{Ag}(\text{NH}_3)_2]^+$ – used in Tollen's reagent

$[\text{Ag}(\text{S}_2\text{O}_3)_2]^{3-}$ – used in photography

$[\text{Ag}(\text{CN})_2]^-$ – forms when silver salts are dissolved in CN^- solutions:

used for electroplating

Transition metal complexes in the body

Haemoglobin is an iron (II) complex. It is surrounded by 4 porphyrin ligands, a globin ligands, and water or oxygen as the sixth ligand.

Cisplatin is an anticancer drug: a 4 co-ordinated complex of Pt^{2+} , ligands are two Cl^- ions and two NH_3 atoms.

Miscellaneous things

In **alkaline solution**, Co (II) and Cr (III) are **oxidised** by H_2O_2 , which acts as an oxidiser or a reducer depending on the conditions.

Co (II) is **oxidised by air** in **ammoniacal solution**.

Organic Chemistry (3.2)

Cracking (thermal) [51]

- **Free radical** reaction
- usual catalyst is **aluminium oxide at 450°C**
- involves **homolytic fission** of C-C bonds only

Main fractions of crude oil [52]

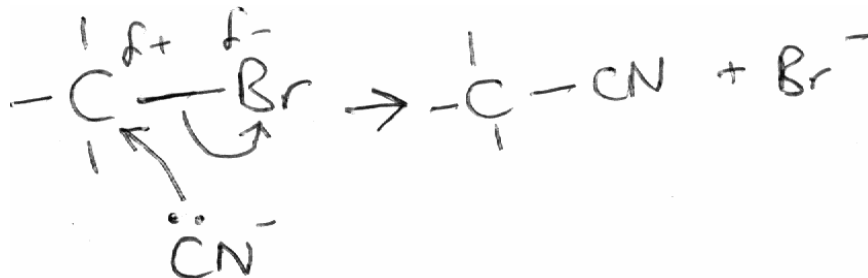
- petrol ($\text{C}_5\text{--C}_{10}$)
- kerosene ($\text{C}_{11}\text{--C}_{12}$)
- diesel ($\text{C}_{13}\text{--C}_{25}$)
- chemical feedstock – Naphtha
- refinery gas
- residue

Pollutants made by cars [53]

- NO_x , CO, unburnt hydrocarbons
- Catalytic converters exist.

Haloalkanes

- undergo **nucleophilic substitution** reactions (susceptible to nucleophilic attack) [61]
- with OH^- ions, and **cyanides** (NaCN , KCN) in aqueous solution:

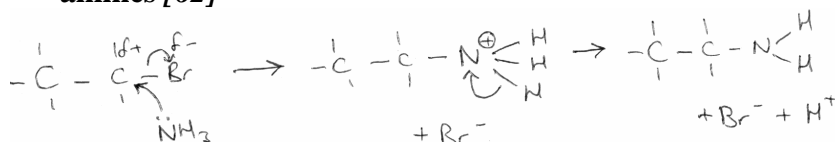


aq. soln
warm

this mechanism also applies to OH⁻ ions.

The end result of the above (bromomethane + NaCN) is **ethanenitrile**, NOT methanenitrile, because the extra C counts as part of the carbon chain.

- with **ammonia**: where the lone pair attacks. **Product: amines [62]**



(haloalkane must be in excess, or secondary amines are formed (multiple substitution))– this makes it less likely that the amine produced will react with Br⁻, also **warm**

The NH₃ reacts with the acid formed (HBr here): NH₃ + HBr → NH₄Br



(they often want this overall equation for the reaction)

Fluoroalkanes are very unreactive (C–F bond v. strong)

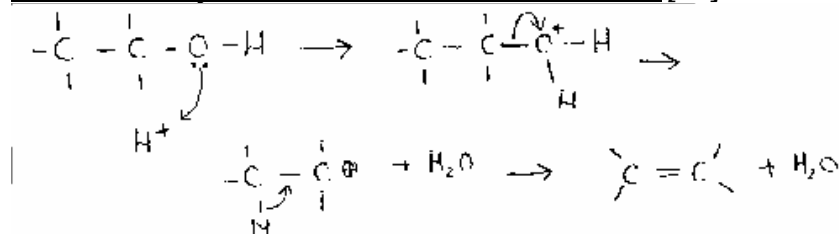
Alcohols [63]

- Made by **hydration of ethene** (phosphoric acid catalyst, 300°C, 65 atm)
- Also by fermenting glucose – impure product

Alcohols: Suitable oxidising agents [64]

- Acidified solution of potassium dichromate (VI) (K₂Cr₂O₇)
 - conditions: **warm** (except for making carboxylic acids from 1° alcohols, in which case **reflux**)
 - Cr₂O₇²⁻ ions change from **orange** to **green** as they are reduced in the reaction.
- 3° alcohols "are not easily oxidised".

Alcohols: Dehydration of alcohols to form alkenes [65]



an **elimination** reaction, reagents conc. sulphuric or conc. phosphoric acid @ **180°C**.

- acid protonates the lone pair on the OH group
- alcohol loses a molecule of water, left with carbonium ion

3. carbonium ion loses proton to produce an alkene.
The acid is actually a catalyst as it is regenerated.

Carbonyl compounds (having a =O but not -OH) [66]

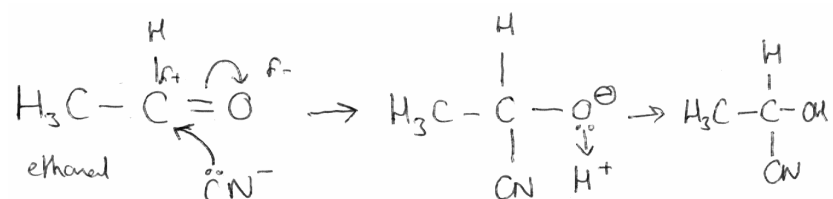
aldehydes/ketones distinguished by

- Tollen's (silver mirror), reagent contains $[\text{Ag}(\text{NH}_3)_2]^+$, this is reduced to silver ppte
Silver mirror formed if an aldehyde is present
- Fehling's solution – normally **deep blue**, goes **brick red** with aldehyde. Contains **copper (II) complex**, that is reduced to **red $\text{CuO}_{2(s)}$** by aldehyde.

[67] Carbonyl compounds may be reduced by (a) NaBH_4 or (b) H_2 gas + Nickel catalyst
symbol for reduction is [H]

- NaBH_4 will **not reduce alkenes**, so it can be used for selective reduction of aldehydes/ketones.
- NaBH_4 is **sodium tetrahydridoborate (III)**
- NaBH_4 works in the **presence of methanol**

Carbonyl reaction with HCN (hydroxynitriles formation) [68]



works for aldehydes and ketones. Reagents: **KCN, dilute sulphuric acid**

The product is a hydroxynitrile, e.g. **2-hydroxypropanenitrile** here.

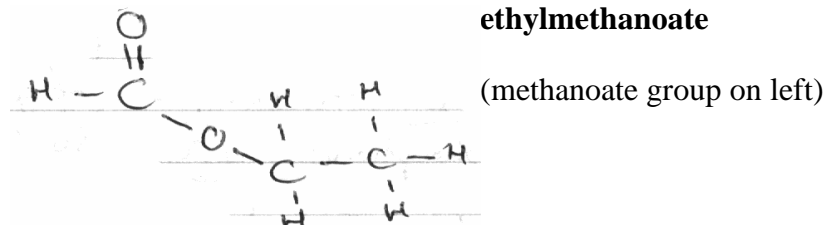
The carbonyl compound actually reacts with HCN gas which is a product of KCN + acid.

Carboxylic acids + esters [69]

acid + alcohol \leftrightarrow ester + water

in the presence of a strong acid catalyst (e.g. sulphuric)

ethylmethanoate



- Esters are used as **solvents, artificial flavourings**, and polymerised to make **synthetic fibres**. [70]
- Natural esters can be **hydrolysed** to other useful products eg. Soap, glycerol, higher fatty acids. [71]

Isomerism [72]

Structural isomer – same molecular formula, different structure

Optical isomer – same structure, not spatially identical (aka enantiomers, stereoisomers). Rotate light in different dir. Also **geometric isomerism** (cis/trans) which are also aka. stereoisomers

Additional [a]

$\text{CH}_4 + \text{Cl}_2$ does not produce any H^\bullet radicals.

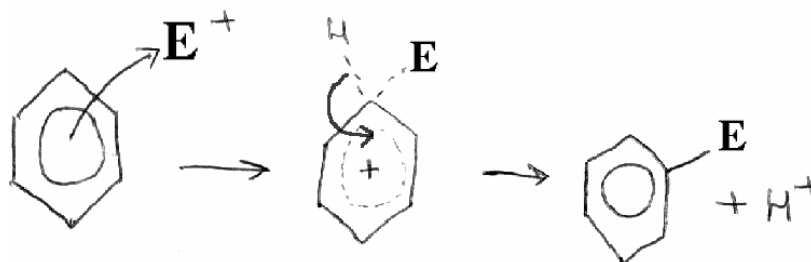
$\text{H}_2 + \text{Cl}_2$ does.

Aromatic Chemistry (6.1)

Benzene – an arene

- bond length is between single and double bond length [59]
- **delocalised, therefore stable** (better than Kekule benzene)

All reactions are **electrophilic substitution**:



A π complex ion (in the middle) is always formed.

Nitration of benzene [60]

Reacts at 50°C with **conc. H_2SO_4** and HNO_3 , sulphuric acid is catalyst.



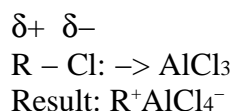
Friedel–Crafts Alkylation/acylation

→ where an alkyl/acyl group is added to a benzene ring.

A **Friedel–Crafts carrier catalyst** is needed, such as AlCl_3 . (Cl may be any halogen)

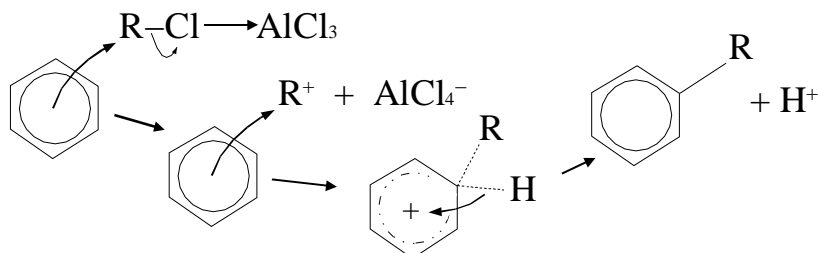
The acyl/alkyl group needs to be bonded to the same halogen, eg. Chloromethane.

The carrier catalyst is a Lewis acid, so it will accept a lone pair from the halogen in the alkyl/acyl group (shown as R here):



That makes the R group much more δ^+ , so that the benzene ring can attack it:

The AlCl_3 is regenerated as the AlCl_4^- is attacked by the H^+ ion to form HCl .

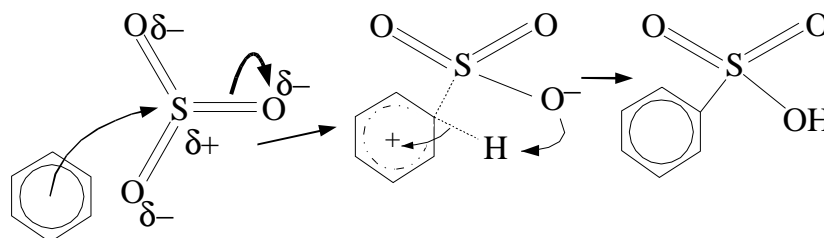


\rightarrow **Ethene** may be used in place of RCl . This reaction is used to make **ethylbenzene** (styrene) used for making **polystyrene**. An additional HCl catalyst is needed, temp. **90°C**

Sulphonation

Very concentrated (**fuming**) sulphuric acid contains the electrophile SO_3 . When

- **refluxed** together for several hours
- mixed with **concentrated HCl at 80°C** , benzene and concentrated sulphuric acid form **benzenesulphonic acid**.



Amines (6.2)

General formula: $\text{R}-\text{NH}_2$

R can be an alkyl or an aryl (benzene) group. If the amine group is the main functional group, then compounds are named as -amine, eg. Phenylamine. If not, then they are named amino- eg. 2-aminopropane.

Named like alcohols. You can have primary, secondary and

tertiary amines.

Amines have base properties. They are lone pair donors.

Methylamine is the best base: the availability of a lone pair is increased by the methyl group, because that group is an **electron releasing** group: it does not pull on the electrons.

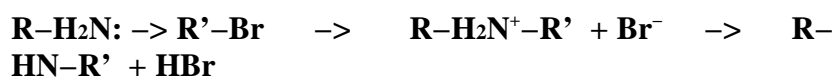
Phenylamine is the worst base: the lone pair on the N is less available, as it is delocalised along with the e^- in the benzene ring.

Ammonia is somewhere between the two, in terms of basic strength.

Reactions of amines

The C–N bond in amines is weak compared to the C–O bond in alcohols, and it's also less polar. Amines tend to react by **nucleophilic substitution**.

eg. reaction with haloalkanes.



(in this reaction it is often important to show what the HBr then does (ie. reacts with the amine) as an overall equation)

The reagent was a primary amine. The product is a secondary amine.

Reaction with haloalkanes adds to the order of the amine:

ammonia + haloalkane \rightarrow 1^o amine

1^o amine + haloalkane \rightarrow 2^o amine

4^o amines exist as salts: quaternary ammonium salts, used to make cationic detergents.

The problem with this reaction is that several products are formed, the product reacts again. This may be limited by putting the haloalkane in excess.

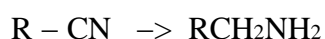
Cationic Detergents from quaternary ammonium salts

usually a long hydrocarbon chain, with a positively charged organic group.

Found in nappy cleaners, hair and fabric conditioners. Have poor cleaning properties but good for germicidal use.

Preparing aliphatic amines

Reduction of a nitrile is one method:–



Reducing agent: H₂ with Ni catalyst (same as reduction (hydrogenation) of ethene to ethane)

Preparing aromatic amines

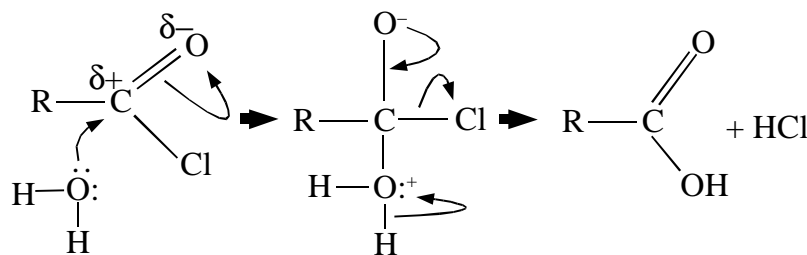
Usually by reduction of an appropriate nitro compound:

eg. **nitrobenzene** + [H] → **phenylamine**

Same reducing agents will work, also tin in HCl.

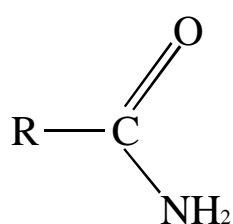
Acid chlorides

These reaction by **nucleophilic addition elimination**, the mechanism's result is similar to substitution, but it's not the same. (Thing to remember: **electrons move to the O on the double bond when the C atom is attacked**)



In the above mechanism, **water** reacts with an **acid chloride** to form a **carboxylic acid**.

If one of the H's in water was replaced with an R' group (making it an alcohol, HOR') the result would be an ester RCOOR'.



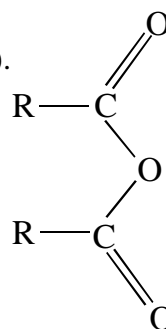
The mechanism will also work for **ammonia**, making RCOONH₂, which is called an **amide** (on left), and also with **amines** (general form NHR') to form RCOONHR': a **substituted amide**.

Reactions of acid anhydrides

This is what an acid anhydride looks like (right).

These are less reactive than acid chlorides.

They exist as **resonance hybrids**: the π electrons can move around between the C-O and C=O bonds.



Acid anhydride + water \rightarrow carboxylic acid

Acid anhydride + alcohol \rightarrow carboxylic acid + ester

Acid anhydride + ammonia \rightarrow carboxylic acid + amide

Acid anhydride + amine \rightarrow carboxylic acid + substituted amide

Acid anhydrides may sometimes be used in place of acid chlorides (both are ethanoylating agents): this is better because they are **cheaper** and **react less vigorously**, and are therefore safer.

Alkene chemistry (6.3)

Note: Alkenes are unsaturated because of the C=C bond

- Alkenes form addition polymers

Inductive effect

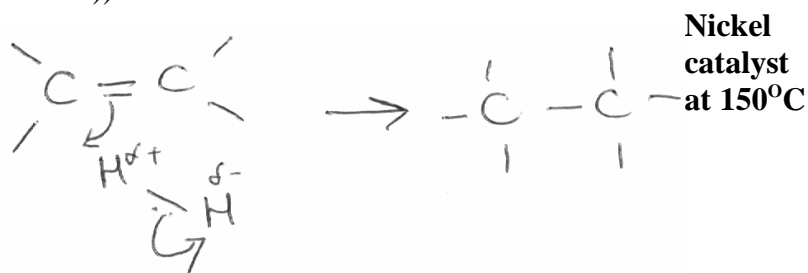
Alkenes react by electrophilic addition. A δ^+ species (eg. H in HBr) adds to the alkene chain on either end of the double bond as that double bond breaks. A carbonium ion is formed at the other end as the bond has broken. (**avoid ever saying C⁺ ion: carbonium is better**)

Markovnikov: The carbonium ion is more likely to form on the end of the bond where it will be surrounded by the most alkyl groups. Alkyl groups exert an **inductive effect; pushing e⁻ towards the carbonium ion**, which makes it **more stable**, and likely to exist for longer, so that the δ^- species will be able to bond to it to complete the reaction.

So, in the reaction between propene and HBr, **2-bromopropane is more likely** to be formed than 1-bromopropane.

Reaction of alkenes with hydrogen

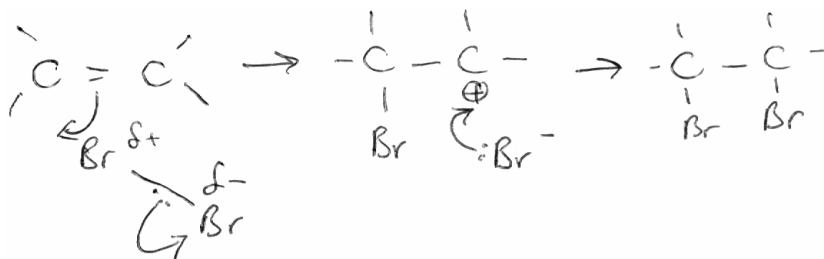
(hydrogenation: adding hydrogen to ethene (margarine reaction))



hydrogenation is NOT hydration. Hydrogenation = adding hydrogen

Reaction of alkenes with bromine

- alkenes decolourise bromine (it adds to them)
- **adding Br₂: (electrophilic addition)**



Happens at room temperature.

- Bromine and hydrogen do not have dipoles, but in the above to reactions, one is **induced** by the presence of the double bond π electrons. **nb. This doesn't work for Chlorine because chlorine's e^- cloud cannot be distorted as easily as Bromine's.** So the reaction **alkene + Cl₂** will not happen
- **HBr** will also add to alkenes. H is the electrophile. Reaction happens in gas phase or in **concentrated aqueous HBr soln.** (I think the aq. bit is important)

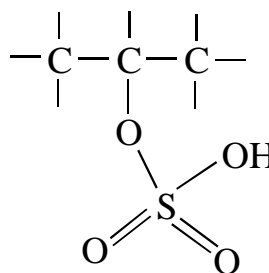
Reaction of alkenes with steam (hydration)

hydration. Alkenes react with **steam** at 300°C with **Phosphoric acid catalyst**,

to form **alcohols**. **PRESSURE:** 65 atm. (this is the industrial way to make alcohols)

Reaction of alkenes with sulphuric acid

Propene reacts with **concentrated sulphuric acid** to form **propyl hydrogen sulphate**. The electrophile is H^+ . Product drawn on right:-



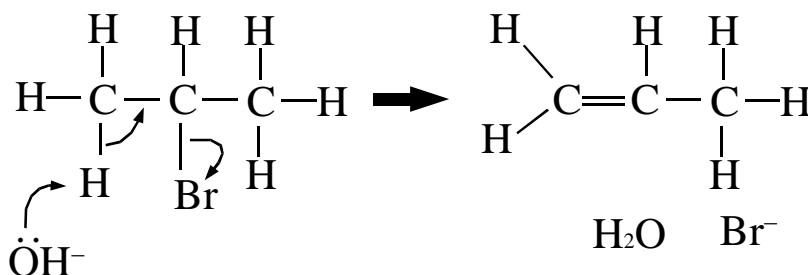
Useful for making alcohols. The product hydrolyses (reacts with water) easily.

Propyl hydrogen sulphate + water \rightarrow propan-2-ol + sulphuric acid

The sulphuric acid is regenerated.

Making alkenes from haloalkanes: **elimination reaction**

In **aqueous solution** haloalkanes react with OH^- ions to form alcohols, as the OH^- ions act as **nucleophiles**. In **ethanolic solution** the OH^- ions act as a **base**, and **alkenes** are formed.



$\text{OH}^-_{(\text{eth})}$ could be produced by dissolving KOH in ethanol. Other products would be KBr, and water.

In this reaction a little propanol will always be formed even though the OH^- ions are not in aqueous solution. The OH^- ions have some nucleophile properties. The reaction type that is favoured depends on:-

- the type of haloalkane (3° favour elimination type)
- base strength of nucleophile: increased by ethanol, reduced by water
- reaction temperature (high temperatures favour elimination)

Epoxyethane

epoxyethane is manufactured by direct synthesis of **ethene** and **oxygen** at **250°C** and **$1\text{--}2\text{MPa}$** with a **silver-based catalyst**. It is a colourless gas, flammable, explosive and very toxic.

It is a very reactive substance due to the strain on the 3 membered ring ($-\text{C}-\text{O}-\text{C}-$). The ring is an area of high electron density, so epoxyethane is **susceptible to electrophilic attack**. Breaking the ring releases much energy.

- It **hydrolyses** (not hydration reaction) to make ethane-1,2-

- diol (antifreeze), usually done at **60°C**.
- It is used to make polymers (terylene)
 - reacts with alcohols to make **alkoxyalcohols** – like diols, but one of the OH groups is an OR group.
 - Alkoxyalcohols used for **solvents in paint industry** and also plasticisers, brake fluid, and detergent manufacture (apparently that's only long chain alcohols though).

Structural Determination

(6.4)

NMR: Nuclei with an odd mass number have a property known as spin.

Variable magnetic field used: nuclei align with the field (α state) or against it (β state). They may resonate from one state to the other when subjected to radio waves. A constant frequency radio transmitter is used, and the radio wave energy absorbed by resonance is detected by a detector coil. Samples rotate at about 25–35 rotations per second to average out any variations in the magnetic field.

The substance being tested is dissolved in tetrachloromethane, for examination to find ^1H nuclei (it contains no ^1H). TMS is mixed with this to provide a reference point for spectrometer calibration. TMS is ideal because:

- the ^1H nuclei are well shielded, so they resonate well above most organic substances
- it gives one sharp easily detected spectrum caused by the 12 equivalent ^1H .
- Non-toxic
- cheap
- non-volatile

Chemical shift

$\delta = (\text{external field at resonance of TMS} - \text{external field at resonance of sample}) / \text{external field at resonance of TMS}$

Chemical shift due to a ^1H nucleus depends mainly on how well it is shielded by e^- density around it. (e^- create a small opposing magnetic field) The H in O–H is less shielded than the H in C–H, because O is more electronegative than C and pulls the e^- away from the H.

...and then, there's singlets, doublets, triplets, quartets and so on. But that's not so hard.

Fingerprint region less than 1500cm^{-1} .

Mechanism notes

- Curly arrows good. Straight arrows bad.
- NMR questions often would like a structure
- Draw lone pairs on everything where $\text{-O:} \rightarrow \text{H}$ etc.
- Don't just say "addition", say "nucleophilic addition" or whatever
- Show the structure when writing out chemical formulae:
 $\text{CH}_3\text{CH}(\text{OH})\text{CH}_3$ (propan-2-ol)
- In equations, it's a good idea to put the catalyst over the \rightarrow

The Reactions and Conditions Table

<u>Reaction</u>	<u>Conditions</u>
Contact Process $2\text{SO}_2 + \text{O}_2$	450°C , a few atm. Pt or V₂O₅ cat.
Haber Process $\text{N}_2 + 3\text{H}_2$	500°C , 180–300 atm. Iron oxide catalysts (some K₂O , and Al₂O₃ act as promoters)
Nitration of benzene	50°C , conc. H₂SO₄ cat
Formation of alcohols/nitriles/amines from haloalkanes	Warm aqueous soln.
Formation of haloalkanes from alkenes	Either: Gas phase or: conc. aqueous soln of HX (X=hal.)
Hydration of alkenes	300°C , 65 atm. H₃PO₄ cat
Hydrogenation of alkenes	H₂ gas at 150°C , with Ni catalyst
Oxidation of alcohols	Warmed gently with K₂Cr₂O₇ Goes from orange to green (Cr³⁺).
Dehydration of alcohols	180°C , conc. H₂SO₄ or H₃PO₄ cat
Reduction of carbonyls	Either: NaBH₄ or: H₂ gas at 150°C , with Ni catalyst
Ethanal + HCN	KCN + dil H₂SO₄ generates HCN in solution for the reaction
Ammonia + O₂ \rightarrow NO	900°C , 7 atm , with Pt cat.

<u>Reaction</u>	<u>Conditions</u>
<i>Formation of ethylbenzene from benzene and ethene</i>	95°C, AlCl₃ cat, in presence of HCl
Formation of alkenes from haloalkanes	Ethanolic soln of KOH. KOH acts as base.
<i>Sulphonation of benzene</i>	80°C, with fuming H₂SO₄
<i>Formation of epoxyethane</i>	250–300°C, 10–20 atm, silver–based catalyst
<i>Formation of ethane–1,2–diol</i>	60°C, H₂SO₄ cat excess of water
<i>Formation of alkoxyalcohols (short chain ones)</i>	Typically 180°C, 10 atm

** end of chemistry **