Chapter 3

Periodicity

The third major topic we tackle is that of periodicity. This is a discussion of the predictable variations and trends in physical and chemical properties of the elements (and the simple compounds of the elements). Its study is interesting because it allows us to make predictions about the behaviour of species of which we have no first hand experience.

3.1 The periodic table

Elements in the periodic table are arranged in order of increasing atomic number, Z. They are *not* arranged in order of increasing atomic mass and you should satisfy yourself that this is the case by looking for examples of this in the table. A group is a vertical column and you should be able to identify groups 1 - 7 and 0. Bear in mind that hydrogen is a very difficult element to place: it commonly forms the H^+ ion but is not metallic, meaning it does not fit neatly into either group 1 or 7. Elements of group 1 (the alkali metals) have 1 outer (valence) electron; group 7 (the halogens) have 7 outer electrons; group 0 (the noble gases) have full outer shells. Elements in the same group tend to have similar chemical and physical properties.

3.2 Physical properties

Physical properties can include aspects such as melting point, size, ionisation energy and electronegativity. You should be able to discuss the variations in these properties as you descend a group (such as group 1 or 7) or as you traverse a period (such as period 3).

3.2.1 Atomic radii

The sizes of atoms can be measured by X-ray diffraction of crystals of an element. It is a judgment call to decide where an atom begins and ends. In order to reduce the margin for error in this atomic radius is commonly measured as half the distance between the nuclei of two bonded atoms of the same element. Of course, if the nucleus is very positive the electrons between the nuclei will 'squash' the electron cloud between them to a certain degree. You might like to be aware that the idea of an atom having a well defined size is somewhat spurious.

As we go down *any* group in the periodic table atomic radii increase. This is because an extra shell of electrons is being added at each period. The outer electron is further from the nucleus by necessity. The extra charge of the nucleus is canceled out by the additional filled shells of electrons (shielding) and the outer electrons experience approximately the same nuclear charge.

As we go across a period the atomic radii fall from left to right. Electrons are being added to the same energy level (n), but the number of protons in the nucleus is increasing, so the positive charge experienced by the outer electrons increases. The shielding effect of the electrons within the same energy level are is very small and the repulsion between the electrons is similarly unable to counteract the effect of increasing the nuclear charge.

3.2.2 Ionic radii

There are two types of ions: cations and anions. Cations are positively charged and anions are negatively charged. **Cations** are *always* smaller than the corresponding neutral atom as a result of increased effective nuclear charge on the outermost electrons. There are two possible explanations for this:

1. The number of energy levels containing electrons may decrease. For instance, when Na^+ is formed from Na there are no longer any electrons in the $n = 3$ energy level, the outermost ones are in $n = 2$. Thus the cation is (much) smaller than the neutral atom.

2. The repulsion between electrons in the same energy level is decreased. If we form Mg^+ from Mg the outermost electron is still in $n = 3$. However, there is one fewer electron in the energy level and so the repulsion between electrons (which has the effect of pushing them away from the nucleus) is lowered and the cation is again smaller than the neutral atom.

Trends in cation size mirror those of atom size.

Anions are *always* larger than their parent atoms. They contain more electrons than protons, but the trends in their properties are the same as cations for the same reasons. Anions are larger than the parent atoms because the effective nuclear charge on the outer electron is decreased when more electrons occupy the same energy level. This also results in increased electron-electron repulsion. In rare cases we may be adding an electron to an otherwise empty shell (consider forming Ne^-) although this is very uncommon.

Almost all anions are larger than almost all cations (the commonest exception to this is F^- and Cs^+).

3.2.3 Ionisation energies

You should be able to define ionisation energy and write an equation for the nth ionisation of any element. You should remember that all ionisation energies are positive as it always requires energy to remove an electron.

Trends in ionisation energies are relatively simple to explain. As we go down a group ionisation energies decrease. We see this to spectacular effect in the reactions of the group 1 metals with water: Li fizzes gently, Cs explodes. The difference is due to the ease of removal of the outer electron. The outer electrons get further from the nucleus as we descend any group. This means that they are easier to remove as they are in higher energy levels. The attraction of the nucleus on the outer electrons is reduced due to the intervening electrons (shielding).

As we go across a period the first ionisation energies increase. This is because the electrons are being removed from the same energy level. Remember that this trend is not a smooth as you might like! The reasons for this trend are the same as the trend down the group. The atoms become smaller from left to right so the effect of the nuclear charge on the outer electrons increases. Shielding is generally ineffective so the effective nuclear charge increases as more protons are added to the nucleus, thus the ionisation energies increase.

3.2.4 Electronegativity

From the above you should be getting the idea that some atoms are better at holding on to electrons than others. *The ability of an atom to attract a shared pair of electrons in covalent bond with another atom* is called **electronegativity**. Electronegativity values are given in the data book but be aware that these are totally arbitrary. They have no units, but most scales agree the same thing: Things towards the top right of the table are more electronegative than the things towards the bottom left.

This should be clear from the discussions of atomic size above. As we descend a group the effect of the nuclear charge on the outer electrons falls as there is increased shielding by the electrons in in between energy levels: electronegativity falls. As we cross a period from left to right the atoms get smaller and the effect of the nuclear charge increases, this means that the atoms are better at attracting negative charge: electronegativity increases. Fluorine is the most electronegative element and francium is the least electronegative. (The top three are F, O and N.)

3.2.5 Melting points

Melting point depends on the structure of the element and the bonding between the particles (atoms or molecules) of the element. There are two types of structure: Giant and simple molecular. Giant structures have higher melting points than simple molecular ones as a rule. Giant structures are split into giant metallic, giant covalent and giant ionic (the last is not considered here as elements only contain metallic or covalent bonds).

Group 1

Li, Na, K, Rb, Cs all have only 1 outer electron. In the metallic structure there is only 1 delocalised electron per positive ion, so the metallic bonds are relatively weak. This means that all group 1 metals have low melting points (for metals). Li has a melting point of 181 °C and Cs has a melting point of only 29 °C (if you held it in your hand it would melt). The melting points fall because the atoms become larger down the group and the strength of the metallic bond falls due to decreasing charge density of the positive ions (the positive charge is spread over a much greater volume).

Group 7

The halogens are all diatomic molecules (HONIFBrCl). Their melting points are: F_2 −220 °C, Cl_2 −101 °C, Br_2 −7 °C, I_2 +114°C. Melting points increase down the group because the forces between the molecules get stronger. These forces are known as *van der Waals* forces (also ‘dispersion’ and ‘London’ forces). Their strength depends on the number of electrons in a molecule - there are 18 electrons in a molecule of F_2 but 106 in I_2. We will cover van der Waals forces later.

Period 3

Na, Mg, Al, Si, P, S, Cl, Ar show a definite trend in melting points. Melting points increase from Na (98°C) to Mg (649°C) to Al (663°C). They all have giant metallic structures and the metallic bonds get stronger as the number of delocalised electrons and charges on the ions increase. Si is a *metalloid*, although it has a giant covalent structure a bit like that of diamond. Covalent bonds are strong and it has a high melting point (1410°C). P (44°C), S (119°C) and Cl (-101°C) all have simple molecular structures - the molecules are held together by van der Waals forces. (Note that the trend is a bit wonky here. This is mainly because S exists as annular S_8 molecules.) Ar is atomic and the van der Waals interactions between these small atoms are very weak, giving it by far the lowest melting point of the period (-189°C)

3.3 Chemical properties

Elements in the same group have similar chemical reactions because of their similar electronic configurations.

Group 1

Read Green & Damji pp91 - 92. All elements in group 1 have 1 outer electron. Reactive metals are stored under oil as they react vigorously with air and water. As the group is descended reactivity increases from Li to Cs. The metals react by losing the outer electron to form an ion with a charge of +1. The reactivity increases because the outer electron is easier to remove as it is further from the positively charged nucleus and therefore not held on as strongly. You should be able to describe the reactions of Li, Na and K with water, chlorine and bromine and write appropriate equations for these reactions. Note that equations in text books may disagree - there are different equations for these reactions in Neuss (p13) and Oxford (p284).

Group 7

Read Green & Damji pp92 - 93. All elements in group 7 have 7 outer electrons. These elements often react by forming negative ions. As the group is descended reactivity decreases from F to I. This can be argued as a function of the increasing size of the atoms (which has the effect of decreasing the electronegativity). You should be able to describe and write equations for the reactions of halogens with halide ions (displacement reactions) and the tests for Cl^-, Br^- and I^- with silver nitrate (with equations).

3.3.1 Change in metallic nature

We shall consider period 3. The s-block elements Na and Mg are clearly metallic. Al has mostly metallic character with some non-metal properties. Si is a metalloid. P, S, Cl and Ar are all non-metals. The electrical conductivity of the elements (given on p 301 of Oxford) is a good illustration of the decrease in metallic character of elements from left to right.

Oxides and chlorides of period 3

The oxides are:
Na_2O, MgO, Al_2O_3, SiO_2, P_4O_6, P_4O_{10}, SO_2, SO_3, Cl_2O and Cl_2O_7
You need to know them. You also need to know the chlorides:
$NaCl$, $MgCl_2$, Al_2O_6, $SiCl_4$, PCl_3, PCl_5, Cl_2

3.4 Period 3 oxides

Physical states

You should be able to explain why the compounds are solids, liquids or gases in terms of their structures. Ionic substances, such as MgO on the left are white solids with high melting points. This is because they have strong ionic bonds which require a large amount of energy to break. SiO_2, which is pure sand, has a giant covalent structure and therefore has a high melting point because the strong covalent bonds require a lot of energy to break. Simple molecular substances on the right such as SO_3 and SO_2 have very low melting points. The weak forces between them (mostly van der Waals forces), called intermolecular forces, are easily overcome with a small amount of energy. *Note that the strong covalent bonds* ***within*** *the molecules are* ***not*** *broken when the substances melt. For instance when* $SO_{2\,(s)} \rightarrow SO_{2\,(l)} \rightarrow SO_{2\,(g)}$ *the* SO_2 *molecules remain intact.* Interestingly, SO_2 is a gas at room temp but SO_3 is a liquid.

Electrical conductivity when molten

The ionic compounds melt to give liquids which conduct electricity. The covalently-bound substances do not. Why?

Acid-base properties

A basic substance will neutralise some acid. The oxides of Na and Mg are basic. Complete the following equations:

$$Na_2O_{(s)} + H_2SO_{4\,(aq)} \longrightarrow$$

$$MgO_{(s)} + HCl_{(aq)} \longrightarrow$$

Aluminium oxide is a bit different. It is ionic, although the electronegativity difference between the elements is small ($3.5 - 1.5 = 2.0$) meaning it has some covalent character. This covalent character is demonstrated by the fact that it is *amphoteric* - it can act as an acid or as a base. The normal behaviour is as a base:

$$Al_2O_{3\,(s)} + 6\,HCl_{(aq)} \longrightarrow$$

It can also act as an acid, as shown by its reaction with sodium hydroxide to form sodium aluminate (or sodium tetrahydroxoaluminate).

$$Al_2O_{3\,(s)} + 2\,NaOH_{(aq)} + 3\,H_2O_{(l)} \longrightarrow 2\,NaAl(OH)_{4\,(aq)}$$

All of the oxides to the right of Al_2O_3 are non-metal oxides and are therefore acidic. $SiO_{2\,(s)}$ is a white solid (it is essentially pure sand) and has a giant covalent structure like that of diamond. It is insoluble in water (why?) and will react with concentrated sodium hydroxide when heated at 350 °C to form sodium silicate:

$$SiO_{2\,(s)} + 2\,NaOH_{(aq)} \longrightarrow Na_2SiO_{3\,(aq)} + H_2O_{(l)}$$

The extreme conditions needed to make this reaction happen show us that while silicon dioxide is acidic, we don't normally notice this property.

The oxides to the right of SiO_2 are all non-metal oxides with simple molecular structures. They are all therefore acidic oxides which react with water forming acids. A typical example would be the reaction of sulphur dioxide gas with water - write the equation and name the acid formed below:

Because these oxides have simple molecular structures they are solids with low melting points, liquids or gases at room temperature.

Reactions with water

Sodium oxide reacts very exothermically with water to give a strongly alkaline solution, which is another piece of evidence that this is a basic oxide. Write an equation for this reaction below:

Magnesium oxide is only slightly soluble in water and so produces a less alkaline solution than Na_2O does (pH is about 10).
Aluminium oxide is insoluble in water. Its bonding has a high percentage of both ionic and covalent characters. Silicon dioxide is also insoluble in water. Why is this?
Phosphorus pentoxide is interesting for historical reasons. Its modern name is phosphorus (V) oxide, which is a lot more appropriate. X-ray crystallography studies in the 1960s showed that the structure of the molecule was actually P_4O_{10} and not P_2O_5. It is a white solid which reacts exothermically and readily with water to produce phosphoric (V) acid. Complete the equation:

$$P_4O_{10\,(s)} + 6\,H_2O_{\,(l)} \longrightarrow$$

Phosphorus (III) oxide, P_4O_6 is a white solid which also reacts with water to produce phosphoric (V) acid. Write the equation for this below.

Sulphur dioxide is the classic non-metal oxide. It is clearly acidic. It is relatively soluble in water and forms sulphurous acid, although the acid itself is difficult to isolate because the reaction is reversible and heating causes it to decompose.

$$SO_{2\,(g)} + H_2O_{\,(l)} \longleftrightarrow H_2SO_{3\,(aq)}$$

Sulphur trioxide reacts with water to form sulphuric acid, and I'm sure you can write an equation for this. With state symbols.

The oxides of chlorine are generally very reactive. They are unstable, often explosive, act as oxidising and bleaching agents. They're really nasty.

$$Cl_2O_{7\,(l)} + H_2O_{(l)} \longrightarrow 2\,HClO_{4\,(aq)} \qquad \text{Perchloric acid}$$

$$Cl_2O_{(g)} + H_2O_{(l)} \longrightarrow 2\,HClO_{(aq)} \qquad \text{Hypochlorous acid}$$

Perchloric acid is a very powerful oxidising agent, hypochlorous acid less so.

3.5 Period 3 chlorides

Physical properties and reactions with water

You should know an unhealthy amount about sodium chloride already. It is the classic ionic solid: White, high melting point, conducts electricity when molten, dissolves in water to give a neutral solution. MgO is similar.
Aluminium chloride has the formula $AlCl_3$. Molecules of $AlCl_3$ only exist in the gas phase, and as molecules they are covalently bound. This gives us yet more evidence that aluminium has significant non-metal character (although it *is* a metal). Aluminium chloride gas can be cooled to give a white solid which is made of dimers with formula Al_2O_6. This is hard to keep because it reacts readily with water (even that in the air) to produce aluminium oxide and hydrochloric acid. There are lots of equations for the reaction of aluminium oxide with water, this is a good un.

$$2\,AlCl_{3\,(s)} + 3\,H_2O_{(l)} \longrightarrow Al_2O_{3\,(s)} + 6\,HCl_{(aq)}$$

You see fumes of HCl evolved during this process because the reaction is very exothermic. The product mixture is very acidic, with a pH of about 0.
$SiCl_4$ is a colourless liquid which reacts with water in a similar way to aluminium chloride. Silicon hydroxide is produced. Complete:

$$SiCl_{4\,(l)} + 4\,H_2O_{(l)} \longrightarrow$$

Fumes of hydrogen chloride gas are also given off. Phosphorus trichloride is similar to silicon tetrachloride. It reacts with water to give both hydrochloric and phosphoric acids (and fumes of hydrogen chloride gas again).

$$PCl_{3\,(l)} + 3\,H_2O_{(l)} \longrightarrow H_3PO_{4\,(aq)} + 3\,HCl_{(aq)}$$

Phosphorus (V) chloride, PCl_5 is a white solid with a simple molecular structure. It reacts readily with water, even in the air, making it difficult to store and dangerous to play with.

$$PCl_{5\ (s)} + 4\,H_2O_{\ (l)} \longrightarrow 5\,HCl_{\ (g\ or\ aq)} + H_3PO_{4\ (aq)}$$

Chlorine gas, Cl_2, is a pale green molecular substance. It dissolves in water to give chlorine water, which is sometimes very pale green in colour. Some of the chlorine **disproportionates** (is simultaneously oxidised and reduced).

$$Cl_{2\ (aq)} + H_2O_{\ (l)} \rightleftharpoons HCl_{\ (aq)} + HClO_{\ (aq)}$$

HClO is called hypochlorous acid. Which is nice.
You should try exercise 3.4 on page 101 of Green & Damji.

3.6 The d-block elements

The d-block elements are also called the transition elements. The first row of transition metals includes the elements from titanium to copper. Scandium ($[Ar]4s^23d^1$) is not a transition metal - it forms only the Sc^{3+} ion which has no d electrons and its compounds are not coloured. Zinc ($[Ar]4s^23d^10$) is not a transition metal for similar reasons - it only forms the Zn^{2+} ion which has a full d subshell (5 orbitals each containing two electrons) and its compounds are not coloured. This means that we need a working definition of what a transition metal is:

A transition element possesses an incomplete d sub-level in one or more of its oxidation states

3.6.1 Variable oxidation states

All transition metals lose 4s electrons before losing their 3d electrons and this means that *all* of them can show an oxidation state of +2. The M^{3+} ion can be formed by loss of a d electron, for instance Fe^{3+} but the M^{4+} ion is rare (MnO_2 being an example). Higher oxidation states normally involve covalent bonding and therefore there are no free ions.

Examples worth knowing

Chromium			
Ox state	+2	+3	+6
Example	$Cr^{2+}\ _{(aq)}$	$Cr^{3+}\ _{(aq)}$	$Cr_2O_7^{2-}\ _{(aq)}$
Colour	blue	green	orange
Manganese			
Ox state	+2	+4	+7
Example	Mn^{2+}	Mn^{4+}	MnO_4^-
Colour	pale pink		deep purple
Iron			
Ox state	+2	+2	
Example	$Fe^{2+}\ _{(aq)}$	$Fe^{3+}\ _{(aq)}$	
Colour	green	brown	
Copper			
Ox state	+1	+2	
Example	$Cu^{+}\ _{(aq)}$	$Cu^{2+}\ _{(aq)}$	
Colour		blue	

3.6.2 Ligands

Ligands are neutral molecules or anions which contain a non-bonding pair of electrons. The lone pairs form a coordinate covalent bond with the metal ion to form a *complex ion*. It is the high charge density (small size, high charge) of the metal ion that attracts ligands.

Types of ligand

Complex formation involves the ‘donation’ of a pair of electrons from the ligand to a cation. The metal ion is, therefore, acting as a Lewis acid, or *electrophile*, whilst the ligand is a Lewis base, or *nucleophile* (meaning ‘electron loving’ and ‘nucleus loving’ respectively). This means that a species *can only act as a ligand if it has a pair of electrons, usually a non-bonded ‘lone’ pair, capable of being donated.* It could be a negative ion such as chloride, oxide or cyanide, or it could be a neutral molecule such as water or ammonia - draw diagrams below to show these species and their lone pairs of electrons.

It should be noted that the tendency of a particular species to complex a metal cation will depend to a large extent on its basic 'strength', which means the ease with which it can donate the pair of electrons. This means that although noble gases possess four lone pairs of electrons in the outer energy level they don't act as ligands because the electrons are too strongly attracted by the nucleus and therefore are not donated to a metal ion. Most of the ligands we will come across are 'monodentate' (one toothed) meaning they only donate one pair of electrons to the metal ion.

Naming of complexes

To name a complex ion we need to

- Work out the overall charge on the complex
- Work out the oxidation state of the central metal species

Convention is to enclose the formula of a complex in square brackets.

$[Fe(H_2O)_6]Cl_2$

This is made up of 2 Cl^- ions and a $[Fe(H_2O)_6]^{2+}$ ion. The water molecules are neutral so the oxidation state of the iron ion must be the same as the charge on the complex ion, which is +2.

$[Cr(OH)_3(H_2O)_3]$

No simple ions are present, so this complex must be electrically neutral. Since each hydroxide ion has a charge of −1, the chromium must be in an oxidation state of +3.

$K_2[CoCl_4]$

The compound consists of two potassium cations $2\,K^+$ and the complex anion $[CoCl_4]^{2-}$. The cobalt must be in an oxidation state of +2 since the charge on each chloride is −1.

The rules for naming compounds are:

1. (a) Electrically neutral complexes such as $[Cr(OH)_3(H_2O)_3]$ have single word names
   (b) In ionic compounds such as $[Fe(H_2O)_6]Cl_2$ and $K_2[CoCl_4]$ the cation is always named before the anion

2. In any complex the ligands are named before the central metal;
    (a) Negative ligands such as Cl^- and OH^- are named before electrically neutral ligands such as H_2O each set being given in alphabetical order.
    (b) The names of negative ligands end in 'o', e.g. Cl^- chloro, OH^- hydroxo, NO_2^- nitrito, CN^- cyano, SO_4^{2-} sulphato
    (c) Neutral ligands usually keep their norma names, except for H_2O aqua, NH_3 ammine, CO carbonyl, NO nitrosyl
    (d) The number of each type of ligand present in indicated by the prefixes: 1 mono-, 2 di-, 3 tri-, 4 tetra-, 5 penta-, 6 hexa-.
3. (a) If the complex is negatively charged, such as $[CoCl_4]^{2-}$, then 'ate' is added after the metal name. Some metals take their Latin name in anionic complexes: Iron (Fe) becomes ferrate, copper (Cu) becomes cuprate, silver (Ag) becomes argentate.
    (b) The oxidation state of the central metal is given in brackets in Roman numerals after its name.

$[\mathbf{Fe(H_2O)_6}]^{2+}\mathbf{2\,Cl^-}$

We name the complex cation first - since we have six water molecules as ligands and iron in an oxidation state of II, this will be

hexaaquairon (II)

The name of the simple anion is then added, in this case chloride. Note that we don't need to say dichloride, since the number of chloride ions needed to balance the charge on the cation is fixed. The full name is therefore

hexaaquairon(II) chloride

$[\mathbf{Cr(OH)_3(H_2O)_3}]$

The complex is electrically neutral and will be named in a single word. We give the negative OH^- ligands before the neutral H_2O, so the name becomes

trihydroxytriaquachromium(III)

$2\,K^+[CoCl_4]^{2-}$

Again we name the cation first, in this case it is simply potassium. The complex anion is tetrachlorocobaltate(II) - note the ending ‘ate’ to indicate the overall negative charge. The full name is therefore

potassium tetrachlorocobaltate(II)

The number of potassium ions is fixed by the overall charge on the anion and need not be specified.

You should be able to describe the shapes of some common complex ions. Those with 6 ligands are *octahedral*, those with 4 are *tetrahedral* ($[CuCl_4]^{2-}$) or *square planar* ($[Cu(NH_3)_4]^{2+}$), those with 2 are *linear*. You should be able to draw a flow diagram for the conversion

$$[Cu(H_2O)_6]^{2+} \longrightarrow [Cu(NH_3)_4(H_2O)]^{2+} \longrightarrow [CuCl_4]^{2-}$$

3.6.3 Colour and magnetic properties of complexes

The origin of colour in transition metal complexes and their behaviour in magnetic fields are most interesting topics and should be mentioned.

Colour

White light is a mixture of all the colours in the visible frequency range of the electromagnetic spectrum. The colours of things in white light depend on whether they absorb all, some or none of the radiation. The colour we observe will be a result of the frequencies which are not absorbed by the sample.
If the sample does not absorb any visible light it will appear white.
If the substance absorbs all frequencies it will be black.

A species will absorb radiation of a particular frequency (colour) only if this corresponds exactly to the energy required to promote one of its electrons from one energy level to a higher one. Colour is normally associated with compounds which have metal ions with an incomplete d sublevel. All free transition metal ions are in fact colourless, so the colours we are familiar with must be due to an interaction between the metal ion and the ligands to which it is attached.

In the free ions all five of the 3d orbitals have the same energy, but when a ligand approaches this arrangement is disturbed and they get split into two sets - some higher, some lower in energy. The difference in energy between the two sets is called the *crystal field splitting energy*, Δ. If an electron can be promoted from the lower set to the upper set by absorbing some coloured light, then the substance will be coloured. If Δ is too large or small then no coloured light will be absorbed and the substance will be white (it absorbs UV or IR radiation).

If the d orbitals are all full or empty then no electrons can be promoted across the energy gap, no visible radiation is absorbed and the substance should be white. In compounds with metals in high oxidation states, such as MnO_4^-, however there can still be colour, but this is caused by charge transfer between the ligands and the metal - in this case we see an intense purple colouration.

3.6.4 Catalytic properties

Hydrogen peroxide decomposition

$$H_2O_2 \xrightarrow{MnO_2} H_2O + \frac{1}{2}O_2$$

The Contact process

Int he manufacture of sulfuric acid, sulfur dioxide is oxidised by air at 450 °C in the presence of a **vanadium (V) oxide** V_2O_5 catalyst

$$2\,SO_{2\,(g)} + O_{2\,(g)} \rightleftharpoons 2\,SO_{3\,(g)}$$

The Haber process

Nitrogen and hydrogen combine under pressure at about 450 °C in the presence of a finely divided **iron** catalyst

$$N_{2\,(g)} + 3\,H_{2\,(g)} \rightleftharpoons 2\,NH_{3\,(g)}$$

Hydrogenation of alkenes

Hydrogenated fats have been in the news recently. Unsaturated vegetable oils can be reacted with hydrogen in the presence of a finely divided **nickel** catalyst in order to increase their melting points (used in the manufacture of margarine). The hydrogen adds to the double bonds in the unsaturated molecules.

Reaction of acidified hydrogen peroxide with iodide

$$H_2O_{2\,(aq)} + 2\,H^+{}_{(aq)} + 2\,I^-{}_{(aq)} \xrightarrow{Fe^{2+}{}_{(aq)}} 2\,H_2O_{(l)} + I_{2\,(s)}$$

via the following mechanism

$$\mathrm{H_2O_2}_{\,\mathrm{(aq)}} + 2\,\mathrm{H^+}_{\,\mathrm{(aq)}} + 2\,\mathrm{Fe^{2+}}_{\,\mathrm{(aq)}} \longrightarrow 2\,\mathrm{H_2O}_{\,\mathrm{(l)}} + 2\,\mathrm{Fe^{3+}}_{\,\mathrm{(aq)}}$$

$$2\,\mathrm{I^-}_{\,\mathrm{(aq)}} + 2\,\mathrm{Fe^{3+}}_{\,\mathrm{(aq)}} \longrightarrow \mathrm{I_2}_{\,\mathrm{(s)}} + 2\,\mathrm{Fe^{2+}}_{\,\mathrm{(aq)}}$$