Grade 12 Chemistry Exam Review

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1 Unit 1

1.1 History of the Atom

1.1.1 Democritus

1. Theorised that all material was constructed out of atoms, tiny particles that could not be further divided.
2. His ideas were laughed at and overshadowed by other, more influential philosophers of his time like Plato and Aristotle.

1.1.2 John Dalton

1. Formalised Democritus's ideas and the concept of the atom.
2. Published the Atomic Theory
    - Atoms are indivisible particles .
    - Elements are composed of atoms. An atom in any given same element will be exactly identical and atoms from different elements will be distinct.
    - Compounds are formed by the joinig of atoms from 2 or more elements.

1.1.3 JJ Thompson

1. Experimentally devised the "Raisin Bun Model" for the atom, which states that the atom was made of many smaller negatively charged particles (Coined "Corpuscles", can the thought of as the "raisins" in the raisin bun) nested in a substance of positive charge (The "bun").
2. The cathode ray tube experiment consisted of shooting the aforementioned "negatively charged particles" through plates with voltage difference. A parabolic path towards the positively charged plate was observed which indicated a constant force was applied to the particle. This force can only exist if the particle was negatively charged, since $|F| = \frac{q|\Delta V|}{|\Delta d|}$, where q is the charge.

1.1.4 Ernest Rutherford

1. Experimentally concluded that the atom contained a dense, positively charged core through the Gold Foil Experiment
2. The gold foil experiment consisted of shooting positively charged α-particles into a gold foil in a room covered in fluorescent paint. The particles bounces off the gold foil and lights up the paint. The "bouncing off" was theorised to be due to the repulsion of the positively charged nucleus and the positively charged particles.

1.1.5 Niels Bohr

1. Theorised that electrons orbit the nucleus in energy levels (orbitals were ”quantised”) through experiments with the electromagnetic spectrum

1.2 The Electromagnetic Spectrum

Electromagnetic waves all have different wavelengths (λ) and travels at the speed of light (c) and has an energy associated with it as follows

$$E = \frac{hc}{\lambda} = hf$$

Due to this energy, every wavelength has, associated with it, unique properties as follows

λ (m)	Classification
$10^2 - 10^4$	Radio
10^{-2}	Micro
10^{-5}	Infrared
$10^{-4} - 10^{-7}$	Visible Light
10^{-8}	Ultraviolet
10^{-10}	X-ray
10^{-12}	γ-ray

Note that For every given wavelength of visible light there exists an associated color as follows, given $n \cdot 10^7$, if $n =$ 4-violet, 5-green, 6-blue, 7-red.

1.3 Spectroscopy

The electromagnetic spectrum is widely used in spectroscopy, 2 main methods of spectroscopy is widely used

1.3.1 Absorption Spectrum

Light is passed through a sample of gas and through a slit. The light passes through the slit and into a prism. The prism then shines the light onto a detector, revealing which wavelengths are left after some are absorbed by the sample of gas.

1.3.2 Emmission Spectrum

Energy (not light) is passed through a sample of gas. Light from the gas is passed through a slit and onto a prism. The prism then shines the light onto a detector, revealing which wavelengths are emmited by the sample.

1.4 Energy Levels

When energy (ex. light) is added shone on a material, the electrons will be excited from ground state and will move up a certain number of energy levels. Afterwards, the electron will drop back down to ground state and in the process, emit a photon. When this is done for different elements, it is found that they have different emission spectra. This implies that certain element’s electrons can only move up a fixed distance, implying the quantisation of energy levels. If the electron is so excited it leaves and never returns, the material is ionised. The energy released by emission corresponds to the distance between the levels the electron is falling through.

1.5 Quantum Model of the Atom

Electrons exist in orbitals in their corresponding energy levels. The orbitals include s, p, d, f, with s being spherical, containing 2 electrons, p being "dumbbell" shaped, containing 6 electrons, and d, f containing 10 and 14 electrons respectively.

1.6 Orbital Box Diagrams

Rules

1. Always fill in every slot with one electron on the way up, then after everything is filled, fill in the second one
2. Electrons in the same spot must have opposite spins (Arrows cannot point the same way)
3. Always fill electrons in order of increasing energy

1.7 Transition Metals and the D-Block

D-Block elements always has possible charge of 2+, and then the possible charges depends on the number of unpaired electrons the element has. The highest possible charge on any given transition metal is given by

$$q = N + 2$$

where q is the charge and N is the number of unpaired electrons the element has.

1.8 Isoelectricity

If 2 or more particles has the same number of electrons, they are said to be isoelectronic, for instance, Cl^- is isoelectronic with Ar since they both have 8 electrons.

1.9 Electron Configuration

A representation of the location of any given element's electrons. Written from the lowest to highest energy level. The format is as follows

$$n\mathrm{Orbital}^N$$

Where n is the energy level and N is the number of electrons in an orbital. For example, fluorine has an electron configuration of $1s^2 2s^2 2p^6 3s^2 3p^5$ since it has 2 electrons in the 1s orbital, 2 electrons in the 2s orbital, 6 electrons in the 2p orbital, and so on.

1.10 VSEPR

The molecular shape and electron distribution of any given compound is determined by the number of electrons and lone pairs it has. The name and angles of molecular geometry is as follows(Note that the notation used AX_nE_m denotes the number of bonds as n and the number of lone pairs as m)

- AX_2: Linear, 180°
- AX_3: Trigonal Planar, 120 °
- AX_2E: Bent, 117°
- AX_4: Tetrahedral, 109.5°
- AX_3E: Trigonal Pyramidal, 107.5°
- AX_2E_2: Bent, 104.5°

- AX_5: Trigonal Bipyramidal, Axial: 90°, Equitorial:120
- AX_4E: See-Saw: Axial:90, Equitorial: 117°
- AX_3E_2: T-Shaped, 90°
- AX_2E_3:Linear, 180°
- AX_6: Octahedral, 90°
- AX_5E: Square Based Pyramidal, 88.5°, 90°
- AX_4E_2: Square BAsed Planar, 90°

The electron distrubution is determined by the AXE formula equivalent to AX_{m+n} for a given molecule, for example, XeF_4: AX_4E2. Its electron distrubution will be $AX_{4+2} = AX_6$, which is octahedral.

1.11 Properties of Matter

1.11.1 Nonpolar Molecules

- Insoluble in water: Since nonpolar molecules only have London Van der Waals (LDVW) forces they cannot replace the H-bonds in water, thus being insoluble.
- Soluble in other nonpolar substances: Since nonpolar molecules only have LDVW they can replace the LDVW in other nonpolar molecules, thus being soluble.
- Low melting and boiling point: Since nonpolar molecules only have LDVW, which is the weakest of all the IMFs, it can be very easily overcome and melt/boil. Larger nonpolar molecules will be harder to melt/boil since there are more LDVW present.
- Soft and Brittle: Polar molecules only have IMFs that can be very easily broken.
- Not Conductive: Nonpolar molecules have neither the ions nor free electrons to be conductive

1.11.2 Polar Molecules

- Soluble in water: Since polar molecules have Dipole-Dipole (DD) forces and some can Hydrogen Bond (HB), they can replace the H-bonds in water, thus being soluble.
- Insoluble in nonpolar substances: Since nonpolar molecules only have LDVW, they cannot replace the LDVW in nonpolar molecules, thus being insoluble.
- Low melting and boiling point: Since polar molecules only have IMFs that can be very easily overcome, they have a low melt/boil point, although still higher than nonpolar molecules (since they have the stronger IMFs).
- Soft and Brittle: Polar molecules only have IMFs that can be very easily broken.
- Not Conductive: Polar molecules have neither the ions nor free electrons to be conductive

1.11.3 Metallic Solids

- Insoluble in water: Since metals are held together by metallic bonds that involves a sea of delocalised electrons, they cannot replace the HB in water
- Malleable and Ductile: Since metals are held together by a sea of delocalised electrons that are free to shift around and move, they are very malleable and ductile.
- High melting and boiling point: Metallic bonds are among the strongest bonds to hold a substance together, and therefore are extremely hard to overcome, making a high melting/boiling point.
- Conductive: The free delocalised electrons can flow to be conductive

1.11.4 Network Solids

- Insoluble in water: Since network solids are held together by a network of pure covalent bonds, they cannot replace the HB in water
- Hard and Brittle: Since network solids are held together by pure covalent bonds that are strong and rigid, they are very hard and brittle.
- High melting and boiling point: Metallic bonds are among the strongest bonds to hold a substance together, and therefore are extremely hard to overcome, making a high melting/boiling point.
- Conductive: Only graphite (3D network solids) is conductive

1.11.5 Ionic Solids

- Soluble in water: Since ionic solids are held together by ionic bonds, they are very polar and can replace the HB in water
- Rigid and Brittle: Since ionic solids are held together in a crystalline structure by ionic bonds, they are very hard. They are also very brittle and can be easy to break if force is applied parallel to the layers in the structure.
- High melting and boiling point: Metallic bonds are among the strongest bonds to hold a substance together, and therefore are extremely hard to overcome, making a high melting/boiling point.
- Conductive: Conductive when aqueous, where ions are free from the structure and able to flow

2 Unit 2

2.1 Collision Theory

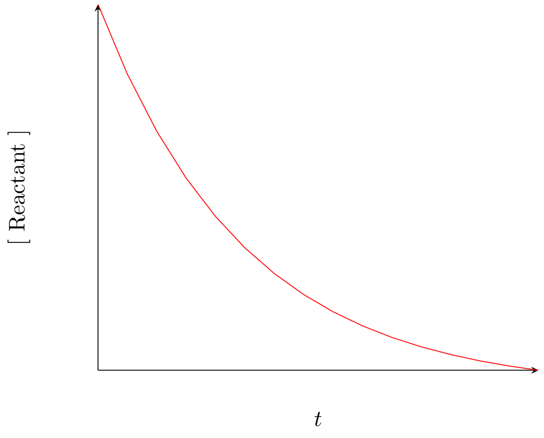
For a reaction to occur the following must be satisfied

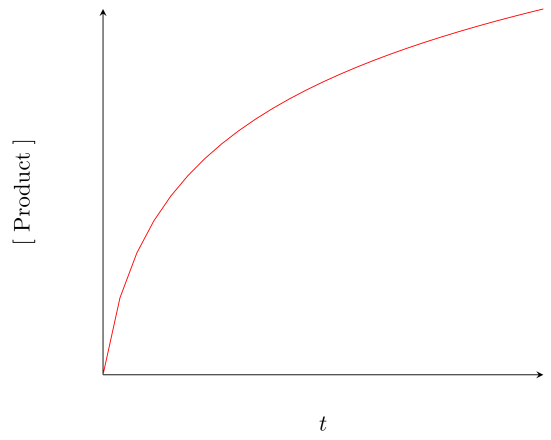
- The particles must collide
- The collision geometry must be favourable
- The collision must have energy greater than the activation energy required for the reaction

And it follows that the rate of reaction will depends on

- The frequency of collisions: The more collisions that happen the faster the reaction goes (Ex. Temperature , concentration of reactant, surface area of particle, pressure of container are all $\propto f$)
- The proportion of effective collisions: More effective collisions occuring means the reaction goes faster (ex. Catalysts lower the activation energy to increase proportion of effective collisons)

In a reaction, the concentration-time graph for any reactant and product will be as follows





This is because when a reaction occurs, at the beginning, there are a lot of reactants around, so it is easier for any given particle of reactant to react with another reactant. This increases the frequency of collisions, thus the rate is high in the beginning. In the end, there are not a lot of reactants around, and the container is mostly occupied by products, decreasing the frequency of collisions, thus the rate slows down.

2.2 Reaction Mechanisms

Reaction mechanisms are the "steps" in which reactions occur. For example, consider the following reaction

$$4HBr + O_2 \rightarrow 2Br_2 + 2H_2O$$

This reactions cannot be done in a single step, so the following mechanisms (steps) are what happens to make this reaction possible

$$\begin{aligned} HBr + O_2 \rightarrow HOOBr &\qquad 50 \text{ kJ} \\ HOOBr \rightarrow 2HOBr &\qquad 20 \text{ kJ} \\ 2(HOBr + HBr \rightarrow 2HOBr) &\qquad 35 \text{ kJ} \\ 4HOBr \rightarrow 2Br_2 + 2H_2O &\quad 5 \text{ kJ} \end{aligned}$$

Notice that some particles are produced in the steps and used up in another. They are known as reaction intermediates. Conversely, if a particle is part of BOTH the reactants and products, it is called a catalyst. Notice again that next to every mechanism, there is an energy associated with it. This energy is the activation energy for each mechanism, and the highest activation energy dictates the overall activation energy of the reaction. The reaction associated with this energy is known as the "rate-limiting" or "slow" step. This step determines the overall rate of reaction. If energy is not given, then the rate-limiting step will be the step with the most colliding particles.

2.2.1 Catalysts

A catalyst works to speed up a reaction by providing an alternate pathway for the reaction to go, usually this pathway has lower activation energy. Thus a catalyst works by reducing the activation energy for a reaction. A catalyst usually speeds up the reaction, but some do slow the rate down. Those are called "inhibitors".

2.2.2 Rules

To represent a general reaction into mechanisms, the following rules must be followed by the mechanisms

- Every mechanism must have less than or equal to 3 particles colliding (Trimolecular collisions)
- Every mechanism must be a plausible single step reaction and must be balanced
- The total effect of the mechanisms must agree with the general reaction (Reaction intermediates must cancel out)

2.3 Quantifying Rates

2.3.1 Rate Law

Let the following general reaction be the slow step of an overall reaction. Remember that for any given step there must not be more than 3 particles colliding, so $\{r_n \geq 0,\ \sum r_n \leq 3\}$

$$r_1R_1 + r_2R_2 + r_3R_3 \rightarrow \text{Products}$$

The rate law for the overall reaction will be as follows

$$r = k[R_1]^{m_1}[R_2]^{m_2}[R_3]^{m_3}$$

Where r is the rate, R_n is the reactant and m_n is the reaction order (It can only be determined experimentally but is usually the coefficients of their respective reactants. If they are 0 then that means their respective reactant has no effect on rate). Note that k is the rate constant and is a result of the proportionality $r \propto [R_1]^{r_1}[R_2]^{r_2}[R_3]^{r_3}$ and that the overall order of the reaction will be simply the sum of the individual orders ($m_{overall} = m_1 + m_2 + m_3$). The rate law cannot contain any reaction intermediates. Should a reaction intermediate ever appear in the rate law, write it in terms of other reactants.

2.3.2 Comparing Rates

The order of reaction only can be found experimentally. In an experiment, the concentration of a reactant must be varied across 2 trials and recorded (c_1 and c_2) and the rate from each trial mut also be recorded (r_1 and r_2). With these information, the order m can be found by rearranging the following equation.

$$\frac{r_2}{r_1} = (\frac{c_2}{c_1})^m$$

For example, consider the following reaction,

$$NH_4^+ + NO_2^- \rightarrow N_2 + 2H_2O$$

The concentration of NH_4^+ was varied across trials 1 and 2. During the first trial, the concentration was 0.1M and the initial rate was found to be $5.4 \cdot 10^{-7}\frac{M}{s}$ and during the second trial, the concentration was 0.2M and the initial rate was found to be $1.08 \cdot 10^{-6}\frac{M}{s}$.

$$\begin{aligned}
\frac{r_2}{r_1} &= (\frac{c_2}{c_1})^m \\
m &= \log_{\frac{c_2}{c_1}} \frac{r_2}{r_1} \\
&= \log_{\frac{0.2}{0.1}} \frac{1.08 \cdot 10^{-6}}{5.4 \cdot 10^{-7}} \\
&= 1
\end{aligned}$$

If the order of NO_2 is also 1 and in the second trial its concentration is 0.2M, the rate constant k can be found as follows. We will use data from the second trial.

$$\begin{aligned}
r &= k[NH_4^+]^1[NO_2]^1 \\
k &= \frac{r}{[NH_4^+][NO_2]} \\
&= \frac{1.08 \cdot 10^{-6}}{(0.2)(0.2)} \\
&= 2.7 \cdot 10^{-5}
\end{aligned}$$

From this information we can write the rate law of the reaction to be $r = 2.7 \cdot 10^{-5}[NH_4^+][NO_2]$

2.4 Thermochemistry

Thermochemistry is the branch of chemistry that is concerned with energy transfer in a chemical system. There are 3 types of chemical systems

- Open Systems: Energy and matter can be transferred to and from surroundings
- Closed Systems: Energy can be transferred to and from surroundings but matter cannot
- Isolated Systems: Energy and matter cannot be transferred to and from surroundings

A chemical reaction can be of 3 types

- Exothermic: Reaction releases energy
- Endothermic: Reaction consumes energy
- Isothermic: Reaction does not release or consume energy

Total energy in a system (E) can be written in terms of kinetic energy (T) and potential energy (U)

$$E = T + U$$

- Kinetic energy is energy associated with motion, and includes vibrational, rotational and translational energy.
- Potential energy is energy stored within or between atoms, and those concerned with chemistry include chemical, nuclear, and electrical potential energy

Thermal energy is a measure of the average kinetic energy of a system, while heat is the amount of energy transferred between substances.

2.5 Heat Transfer

When heat is supplied to an object, its average kinetic energy will increase. It is described by the following equation

$$\begin{aligned} Q &= mc\Delta T \\ &= mv(T_f - T_i) \\ T_f &= \frac{Q}{mv} + T_i \end{aligned}$$

This equation can be used to solve for the final temperature of an object when heat Q is supplied to it. When heat is transferred between objects, the object with the lower temperature will have increased kinetic energy while the object with the higher temperature will have decreased kinetic energy. This change will continue happening until a final temperature is reached. And can be described by a modified version of the above equation. Let object 1 be the one losing energy and object 2 be the one gaining energy (object 1 must start hotter than 2).

$$\begin{aligned} -Q_1 &= Q_2 \\ -m_1c_1(T_f - T_{1i}) &= m_2c_2(T_f - T_{2i}) \\ T_f &= \frac{T_{1i}m_1c_1 + T_{2i}m_2c_2}{m_2c_2 + m_1c_1} \end{aligned}$$

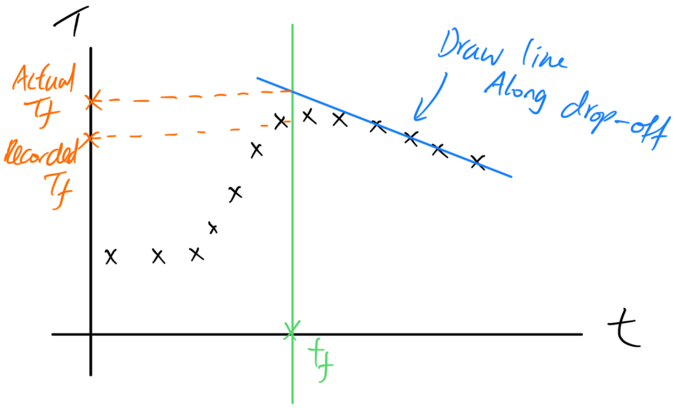
During state change, there will be energy transfer but no temperature change. This concept is not tested in the exam.

2.6 Calorimetry

A calorimeter measures how much energy is in a substance by combusting it and see how much the surrounding water heats up. A typical calorimeter includes a box of water with a lid, followed by a thermometer and stirrer in the water. In the water, the sample is combusted in a closed box where energy can be transferred into the water. In a bomb calorimeter, an ignition plug is attatched to the sample to ignite the combustion. In calorimetry, the following assumptions are made

- The sample transfers all energy to the water
- The calorimeter is perfectly isolated from the surroundings
- The calorimeter does not absorb or release energy
- The sample solution has the same specific heat capacity as water

If the calorimeter does transfer energy to the surroundings, the following steps in analysis could be used to correct this innacuracy



2.7 Enthalpy

Enthalpy is the total amount of energy in a system, and is denoted by ΔH. There is no absolute entlapy, as it is relative. So therefore in all instances it must always the "change in enthalpy", ΔH.

2.7.1 Enthalpy in calorimetry

In calorimetry, enthalpy can be determined by the following equation

$$\Delta H = -Q_{H_2O}$$

For example, in a calorimeter where 2g of KCl is dissolved and combusted in 100g of H_2O at room temperature. The final temperature is 50 degrees celcius. The enthalpy can be calculated as follows

$$\begin{aligned} Q_{H_2O} &= m_{H_2O} c_{H_2O} (T_f - T_{1H_2O}) \\ &= (100)(4.18)(50 - 25) \\ &= 10450 \text{ J} \\ \Delta H &= -Q_{H_2O} \\ &= -10450 \text{ J} \end{aligned}$$

And the molar enthalpy can simply be found by dividing the enthalpy by the number of moles of KCl.

2.8 Hess's Law

When given a reaction with all it's mechanisms, the individual mechanisms must be scaled in order to fit the overall reaction. After doing so, adding up all the individual enthaplpies will yield the net enthalpy. For example.

$$CaO + H_2O \to Ca(OH)_2$$

Can be split into the following mechanisms

$$\begin{aligned} Ca + \frac{1}{2}O_2 &\to CaO \qquad \Delta H = -635\text{kJ} \\ H_2 + \frac{1}{2}O_2 &\to H_2O \qquad \Delta H = -286\text{kJ} \\ Ca + H_2 + O_2 &\to Ca(OH)_2\ \Delta H = -982\text{kJ} \end{aligned}$$

If we added them all up now, it would look nothing like the original equaton. We first notice that the CaO of the first equation is on the wrong side, so we need to flip the equation. We can do so my multiplying the first equation by -1. This yields

$$CaO \to Ca + \frac{1}{2}O_2 \qquad \Delta H = +635\text{kJ}$$

We then notice that the H_2O is also on the wrong side, so we also need to flip the second equation by multiplying it by -1. This gives

$$H_2O \to H_2 + \frac{1}{2}O_2 \qquad \Delta H = +286\text{kJ}$$

The third equation has nothing wrong. By inspection, the hydrogens and the oxygens should cancel out when added together with the other equations, and the $Ca(OH)_2$ is on the correct side, so we do not need to do anything to it except multiply it by 1. This gives

$$Ca + H_2 + O_2 \to Ca(OH)_2\ \Delta H = -982\text{kJ}$$

Adding all the modified equations together gives the following

$$\begin{aligned} CaO + H_2O + Ca + H_2 + O_2 &\to Ca + \frac{1}{2}O_2 + H_2 + \frac{1}{2}O_2 + Ca(OH)_2 \\ CaO + H_2O + Ca + H_2 + O_2 &\to Ca + O_2 + H_2 + Ca(OH)_2 \end{aligned}$$

Cancelling out the particles that appears on both sides gives

$$CaO + H_2O \to Ca(OH)_2$$

Which is precisely the original equation. Knowing that this equation is correct, we can add together the modified enthalpies.

$$\begin{aligned} \Delta H &= +635\text{kJ} + 286\text{kJ} - 982\text{kJ} \\ &= -61\text{kJ} \end{aligned}$$

2.9 Enthalpies of Formation and Bond Enthalpies

Enthalpies of formation are average enthalpies for the formation of each molecule. Bond enthalpies are the energy required to break or form one mole each bond. The bond enthalpy of the forming of one mole of a molecule is simply the sum of bond enthalpies of the reactant that is used to make the molecule substracted from the sum of bond enthalpies of the molecule (the product) itself . They are not very accurate since they are only an average, and enviromental conditions may affect these numbers heavily.

3 Unit 3

3.1 Introduction to Equilibrium

For any given reaction, there exists an equilibrium, where the rate of the forward reaction is equal to be rate of the reverse reaction, thus no visible changes in concentration can be seen. This equilibrium can be found on the [ ]-t graph as the "flatline" region.

3.2 La Chatalier's Principle

Several factors can affect the equilibrium of any given reaction. Note that "shift left" denotes an increase of reactants and "shift right" denotes an increase in products

3.2.1 [Reactant] increase

The concentration of the added reactant will increase sharply, followed by a smooth decrease in all [Reactants] and a simultaneous smooth increase in all [Products]

3.2.2 [Products] increase

The concentration of the added product will increase sharply, followed by a smooth decrease in all [Products] and a simultaneous smooth increase in all [Reactants]

3.3 Temperature Increase

The concentrations of whichever side energy is on will decrease smoothly, while the side with no energy term will have their concentrations increased.

3.4 Pressure Increase

The concentrations of all reactants and products increases, followed by a smooth increase in concentration in the side with less gaseous particles, and a decrease in concentration in the side with more gasous particles. Note that a decrease in volume is equal to an increase in pressure.

3.5 Equilibrium Constant

Consider the following general reaction, where R_n and P_n are reactants and products in the forward reaction respectively and r_n and p_n are their respective coefficents

$$r_1R_1 + r_2R_2 + ... + r_nR_n \leftrightarrows p_1P_1 + p_2P_2 + ... + p_nP_n$$

For every reaction, there will exist an equilibrium, and the equilibrium concentrations of every reactant and product can be expressed in a single number - the equilibrium constant. The equilibrium constant K for any given reaction in the above form can be calculated as follows

$$K = \frac{[P_1]^{p_1}[P_2]^{p_2}...[P_n]^{p_n}}{[R_1]^{r_1}[R_2]^{r_2}...[R_n]^{r_n}} = \frac{\prod[P_n]^{p_n}}{\prod[R_n]^{r_n}}$$

For example, consider the following reaction

$$2CO_2 \leftrightarrows 2CO + O_2$$

If the equilibrium concentrations are $[CO_2]$=0.25M, $[O_2]$=2.2·10^{-3}M,$[CO]$=4.3·10^{-3}M, the equilibrium constant would be

$$\begin{aligned} K &= \frac{[CO]^2[O_2]}{[CO_2]^2} \\ &= \frac{(4.3 \cdot 10^{-3})^2(2.2 \cdot 10^{-3})}{(0.25)^2} \\ &= 6.4 \cdot 10^{-7} \end{aligned}$$

3.6 Reaction Quotient

For every reaction, the concentrations of the reactants and products in it's current state can be represented by a reaction quotient (Q), which works the exact same way as the equilibrium constant, except the concentrations are current concentrations instead of equilbirium concentrations, and the system is said to be in equilibrium if $Q = K$. Therefore it can be represented as follows for the same general reaction in a non-equilibrium state

$$Q = \frac{[P_1]^{p_1}[P_2]^{p_2}...[P_n]^{p_n}}{[R_1]^{r_1}[R_2]^{r_2}...[R_n]^{r_n}} = \frac{\prod[P_n]^{p_n}}{\prod[R_n]^{r_n}}$$

For example, consider the following reaction

$$2CO_2 \leftrightarrows 2CO + O_2$$

If the current state of the reaction are $[CO_2]$=0.20M, $[O_2]$=3.2·10^{-3}M,$[CO]$=1.3·10^{-3}M, the equilibrium constant would be

$$\begin{aligned} Q &= \frac{[CO]^2[O_2]}{[CO_2]^2} \\ &= \frac{(1.3 \cdot 10^{-3})^2(3.2 \cdot 10^{-3})}{(0.20)^2} \\ &= 1.4 \cdot 10^{7} \end{aligned}$$

In general, if $Q < K$, the system will want to increase Q to reach equilibrium, so the system will shift towards the product side, and if $Q > K$, the system will want to decrease Q to reach equilibrium, so the system will shift towards the reactant side.

3.7 Equilibrium Calculations

Find the equilibrium concentration for the following reaction, $K = 4.5$ and the starting concentrations are $[NO_2] = 0.25M$, $[N_2O_4] = 0M$

$$N_2O_4 \leftrightarrows 2NO_2$$

We first see that $Q > K$, since there are no reactants, $Q \to \infty$, so the reaction must shift left, therefore we implement a change variable ϕ. The coefficients of ϕ will correspond to the coefficient of its respective reactant/product, and the sign of ϕ will correspond to the shift (Shift to opposite side, sign is negative, shift to same side, sign is positive)

	N_2O_4	NO_2
I	0M	0.25M
C	$+\phi$	-2ϕ
E	$+\phi$	$0.25 - 2\phi$

From the E row, we make a new reaction quotient and equate it with K to find ϕ

$$\begin{aligned} K &= \frac{[NO_2]^2}{[N_2O_4]} \\ 4.5 &= \frac{(0.25-2\phi)^2}{\phi} \\ 4.5\phi &= \frac{(0.25-2\phi)^2}{\phi} \\ \phi &= 0.1 \end{aligned}$$

Note that for every polynomial, there may be multiple roots, but only one is valid, so the other one will be dismissed. Finally, we can see that

$$\begin{aligned} [NO_2] &= 0.25 - 2\phi = 1.3M \\ [N_2O_4] &= \phi = 0.1M \end{aligned}$$

3.8 Acids and Bases

- In Arrhenius theory, an acid dissociates in water to release hydrogen and a base dissociates in water to release hydroxide.
- In Bronstead-Lowry Theory, an acid ionises in water to donate protons and a base ionise in water to accept protons.
- Strong acids and bases ionise fully while weak acids and bases only partially ionise

3.9 pH and pOH

pH and pOH are a measure of the concentration of H^+ and OH^- ions respectively, and are defined as follows

$$\begin{aligned} pH &= -\log[H^+] \\ pOH &= -\log[OH^-] \end{aligned}$$

In water, the following reaction is constantly at equilibrium

$$2H_2O \leftrightarrows H^+ + OH^-$$

And we can write the equilibrium constant as follows

$$K = [H^+][OH^-]$$

Since there is no concentration of water, the water term is omitted. In water at SATP, the value for K is $1.00 \cdot 10^{-14}$. This value is knwon as K_W. Substituting this value and then taking the negative log of both sides gives

$$\begin{aligned} 1.00 \cdot 10^{-14} &= [H^+][OH^-] \\ 14 &= -log[H^+] - log[OH^-] \\ 14 &= pH + pOH \end{aligned}$$

And thus yields the relationship between pH and pOH. For example, in a solution of 2.0M NaOH, the pH can be easily found as follows

$$\begin{aligned} [NaOH] &= [OH^-] = 2.0M \\ 14 &= pH + pOH \\ pH &= pOH - 14 \\ &= 14 - \log{[OH^-]} \\ &= 14 - \log 2.0 \\ &= 13.70 \end{aligned}$$

Note that in base-10 logarithms, the integer part of the solution are not significant, so 2 sig-figs will mean 2 decimal places.

3.10 Acid and Base Constants

For weak acids and bases, finding pH and pOH is not that straightfoward. Consider the following acid ionisation reaction

$$HA + H_2O \leftrightarrows A^- + H_3O^+$$

The equilibrium constant (We call K_A this time) can be easily found

$$K_A = \frac{[A^-][H_3O^+]}{[HA]}$$

This value is called the acid constant, and is unique for every weak acid. In most cases, the K_A is given, and the pH is found by simply rearranging for $[H_3O^+]$, since $[H_3O^+] = [H^+]$. For example, consider the following reaction. Choose $[HCN] = 0.25M$ and $K_A = 4.9 \cdot 10^{-10}$

$$HCN + H_2O \leftrightarrows CN^- + H_3O^+$$

We begin by constructing an ICE table

	$[HCN]$	$[CN^-]$	$[H_3O^+]$
I	0.25M	0	0
C	$-\phi$	$+\phi$	$+\phi$
E	$0.25-\phi$	$+\phi$	$+\phi$

We notice that the concentration of hydrocyanic acid is far greater than the K_A, and the ratio between them far exceed 500, therefore we can say that $0.25-\phi \approx 0.25$. Now we can write the quotient

$$\begin{aligned} 4.9 \cdot 10^{-10} &= \frac{\phi^2}{0.25} \\ \phi &= \sqrt{4.9 \cdot 10^{-10} \cdot 0.25} \\ &= 1.1 \cdot 10^{-5} \end{aligned}$$

Plugging this back into the concentration and pH equations we can easily find the pH

$$\begin{aligned} [H_3O^+] &= \phi = 1.1 \cdot 10^{-5} \\ pH &= -\log[H_3O^+] \\ &= -\log 1.1 \cdot 10^{-5} = 4.96 \end{aligned}$$

For every basic ionisation reaction

$$B + H_2O \leftrightarrows HA^+ + OH^-$$

The equilibrium constant (We call K_B this time) can be easily found

$$K_B = \frac{[HB^+][OH^-]}{[B]}$$

This value is called the basic constant, and is unique for every weak base. In most cases, the K_B is given, and the pH is found by simply rearranging for $[OH^-]$ and putting it into the equation that relates it with pH. For example, consider the following reaction. Choose $[NH_3] = 0.25M$ and $K_B = 1.8 \cdot 10^{-5}$

$$NH_3 + H_2O \leftrightarrows NH_4^+ + OH^-$$

The ICE table is as follows

	$[NH_3]$	$[NH_4^+]$	$[OH^-]$
I	0.25M	0	0
C	$-\phi$	$+\phi$	$+\phi$
E	$0.25-\phi$	$+\phi$	$+\phi$

We notice that the concentration of ammonia is far greater than the K_B, and the ratio between them far exceed 500, therefore we can say that $0.25-\phi \approx 0.25$. Now we can write the quotient

$$\begin{aligned} 1.8 \cdot 10^{-5} &= \frac{\phi^2}{0.25} \\ \phi &= \sqrt{1.8 \cdot 10^{-10} \cdot 0.25} \\ &= 6.7 \cdot 10^{-6} \end{aligned}$$

Plugging this back into the concentration and pH equations we can easily find the pH

$$\begin{aligned} [OH^-] &= \phi = 6.7 \cdot 10^{-6} \\ pH &= 14 - pOH \\ &= 14 + \log[OH^-] \\ &= 14 + \log\left(6.7 \cdot 10^{-6}\right) = 8.83 \end{aligned}$$

Note that there exists a relationship between an acid's K_A and its conjugate base's K_B, and vice versa, as follows

$$K_W = K_A K_B$$

This equation implies that the stronger an acid is, the weaker it's conjugate base will be, and vice versa. This rule also applies to bases and their conjugate acids.

3.11 Buffer Solutions

A buffer solution is made with a weak acid/base with its respective ion in an aqueous solution. Buffer solutions will work to resist pH changes. A general buffer solution of an acid can be represented as follows

$$HA + H_2O \leftrightarrows A^- + H_3O^+$$

Buffer solutions are only effective if the pH is in an effective range, in which $pH = pK_A \pm 1$. The pH of an acid buffer solution can be found by the Henderson-Hasselbach equation, which is as follows

$$pH = pK_A + \frac{[A^-]}{[HA]}$$

The buffer solution of a base can be represented as follows

$$B + H_2O \leftrightarrows HA^+ + OH^-$$

And the Henderson-Hasselbach equation can be modified to find the pH of a base as follows

$$pH = pK_A + \frac{[B]}{[HB^+]}$$

Where pK_A is the pK_A of its conjugate acid. For example, consider a buffer made of $0.10MCH_3COOH$ and 5g CH_3COONa in a 1L container

$$\begin{aligned}
[CH_3COO^-] &= [CH_3COONa] \\
&= \frac{m}{(2Ar(C) + 3Ar(H) + 2Ar(O) + Ar(Na))V} \\
&= 6.09 \cdot 10^{-2}M \\
pH &= pK_A + \frac{[CH_3COO^-]}{[CH_3COOH]} \\
&= 4.74 + \log\frac{6.09 \cdot 10^{-2}M}{0.10M} \\
&= 4.53
\end{aligned}$$

When $[OH^-]$ is added to a acid buffer solution, the system will seek to compesate by shifting right to produce more $[H_3O^+]$, whereas when $[H_3O^+]$ or $[H^+]$ is added to a acid buffer solution, the system will seek to compensate by shifting left to produce less $[H_3O^+]$. This effect is the opposite when in a basic buffer solution. where when $[OH^-]$ is added to a basic buffer solution, the system will seek to compesate by shifting left to produce less $[OH^-]$, whereas when $[H_3O^+]$ or $[H^+]$ is added to a basic buffer solution, the system will seek to compensate by shifting right to produce more $[OH^-]$ to balance out the pH change. Calculations like these are no different than normal ICE table calculations, except the change is the $[H_3O^+]$ or $[OH^-]$ added and the sign is dependent on the effects mentioned above.

3.12 Titration Curves

Titration curves are a combination of 3 calculations - pH_1, pH_f and V_{eq}. Where pH_1 is simply the pH of the sample, pH_f is the pH of the titrant, and V_{eq} is the volume calculated by using the equation

$$V_{eq} = \frac{c_{sample}V_{sample}}{c_{titrant}}$$

The equilibrium pH will change depending on how strong the conjugate acids and bases of the sample and titrant are. If conjugate base is stronger than conjugate acid, equilibrium pH¿7, and vice versa. If they are equal, equilibrium pH=7.

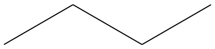
4 Unit 4

4.1 Hydrocarbons

Every organic compound contains carbon because carbon has a small atomic radius and can form up to 4 bonds. Hydrocarbons are made of hydrogen and carbon and are split into 2 categories: Aromatic and aliphatic. Aromatic hydrocarbons contains benzene rings while aliphatic ones do not.

4.2 Alkanes

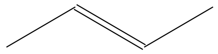
Alkanes are made of chains of CH_2 single bonded with each other with a CH_3 at each end. Alkanes are named with a number prefix (meth-, eth-, etc.) depending on the number of carbons in the chain, and are named with an -ane ending. For example, consider the following



It has 4 carbons and thus has the ”but-” prefix. So it is called butane. Note that in a line diagram every node represents a carbon

4.3 Alkenes

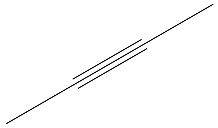
Alkenes are hydrocarbons with double bonds. Alkenes are named like alkanes but with a number denoting where the double bonds are followed by the "-ene" ending. If there are more than one double bond, list all of them in ascending order followed by "-ene" with a number prefix (ex. di, tri, etc.). For example,



There is a double bond in the second carbon, and there are 4 carbons in total, so its but-2-ene

4.4 Alkynes

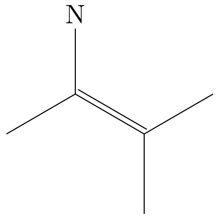
Alkynes are hydrocarbons with triple bonds. They are named like alkenes except the "-ene" ending is changed to "-yne". For example



Theres a triple bond in the second carbon, and there are 4 carbons in total, so its but-2-yne

4.5 Side Chains and Other Naming Rules

To name a side chain, simply put the number of the side chain's carbon and then the side chain with the "-yl" suffix (ex. methyl, ethyl, etc.). Special names include the iodine side chain: iodo-, chlorine side chain: chloro-, the fluorine side chain: fluoro-, the nitrogen side chain: nitro, the benzene side chain, : phenyl-, and the isopropyl side chain: isopropo-. It is generally recommended to avoid using the isopropyl side chain and only save it as a last resort. Names should be ordered in alphabetical order and alkynes should be given priority over alkenes which has priority over alkanes when numbering carbons. If there are no double or triple bonds, the object that is closest to the end has priority. An example of naming is as follows



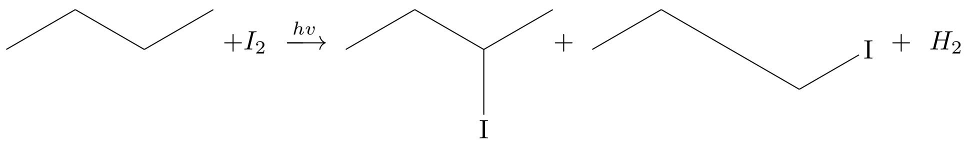
The main chain has 4 carbons, there is a methyl group on the second carbon and a nitro group on the third. There is also a double bond on the second carbon. So this is 2-methyl-3-nitrobut-2-ene.

4.6 Properties of Hydrocarbons

- Insoluble in water: Since nonpolar molecules only have London Van der Waals (LDVW) forces they cannot replace the H-bonds in water, thus being insoluble.
- Soluble in other nonpolar substances: Since nonpolar molecules only have LDVW they can replace the LDVW in other nonpolar molecules, thus being soluble.
- Low melting and boiling point: Since nonpolar molecules only have LDVW, which is the weakest of all the IMFs, it can be very easily overcome and melt/boil. Larger nonpolar molecules will be harder to melt/boil since there are more LDVW present. Note that Alkynes has a higher boiling point than alkanes which has a higher boiling point than alkenes. This corresponds to how well they pack with each other. Since better packing means that they are closer together and thus the LDF are stronger.
- Unreactive: Unreactive since it is hard to break a C-C and C-H bond. Alkynes are more reactive than alkenes which are more reactive than alkanes. This is because it is easier to convert a double or triple bond into a single or double bond respectively than it is to break a single bond.

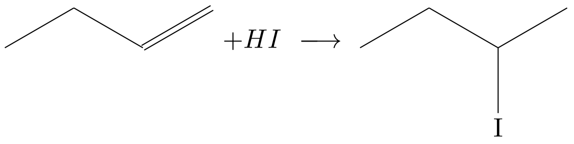
4.7 Reactions of Hydrocarbons

4.7.1 Halogenation



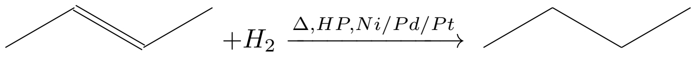
All halogenation reactions need hv (light energy) unless the halogen is fluorine.

4.7.2 Hydrohalogenation



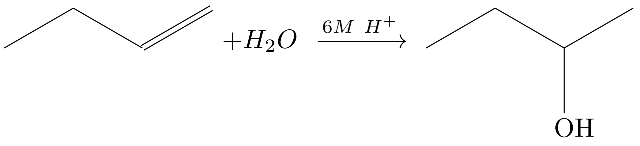
Note that in hydrohalogenation, Markovnikov's rule apply, meaning that the hydrogen will always take the outer carbon (the carbon with the most hydrogen already attatched to it).

4.7.3 Hydrogenation



Note that in hydrogenation, only one of the 3 metal catalysts is needed, but in exam it is required to write all 3.

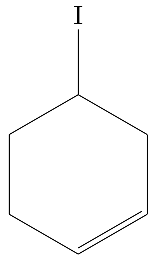
4.7.4 Hydration



Note that in hydration, Markovnikov's rule still applies as the H_2O seperate into H and OH and forms a secondary alcohol.

4.8 Cyclic Hydrocarbons

Cyclic hydrocarbons are hydrocarbons that forms a ring. They are names like alkanes and alkenes (note that the numbering must "go through" the double bond should there be any) except with the "cyclo-" prefix. Small molecules like cyclopropane are prone to breaking since they have a lot of steric strain. Cyclic hydrocarbons cannot contain triple bonds since then the ring will be too big. The following is an example in naming.

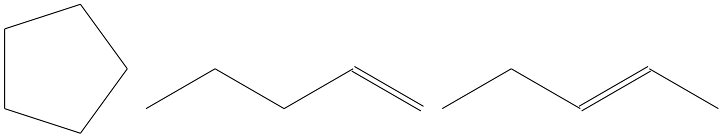


Numbering through the double bond, the double bond is at 1 and an iodo group at 4, so it is 4-iodocyclohex-1-ene.

4.9 Isomers

4.9.1 Structural Isomers

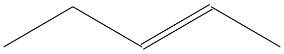
Structural isomers are molecules with the same formula but with different structures, for example, some structural isomers of C_5H_{10} include



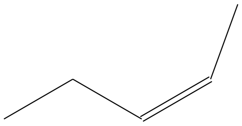
These molecules may have the same chemical formula, but they have vastly diffrerent properties so it is important to distinguish between them.

4.9.2 Geometric Isomers

Geometric isomers are molecules that only differ by rotation about a C=C bond. For example,



and



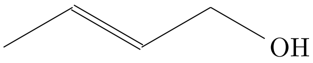
These two molecules differ only because the methyl group is on opposite sides. The left one is the E-Isomer and the right one is the Z-Isomer.

4.9.3 E-Z Priorities

E-Z Priorities are assigned based on which of the 2 atoms bonded to the C=C has more protons. The higher the proton count, the higher the priority. If in an isomer, the 2 prioritised atoms are on the same side, the isomer is known as the Z-Isomer. If they are on different sides, It is the E-Isomer.

4.10 Alcohols

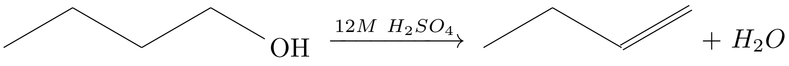
Alcohols are hydrocarbons with the functional group $-OH$ attatched as the highest priority group. Since they have the -OH group, they can HB with water and are thus very soluble. Alcohols are granted the highest numbering priority. Alcohols are named with the ”-ol” suffix. For example



This has a double bond at 2 and an OH group at one. This has 4 carbons so its but-2-en-1-ol

4.11 Reactions of Alcohols

Alcohols can undergo elimination to return to an alkene as follows



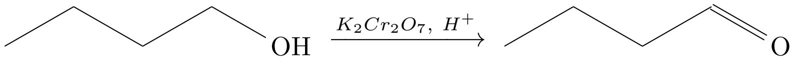
Alcohols can also be oxidised to become either an aldehyde or a ketone, which will be discussed later.

4.11.1 Types of Alcohols

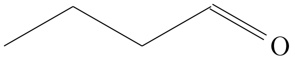
- 1° Alcohol: $-OH$ attatched to CH_2, can react twice and will only react in the oxidation test.
- 2° Alcohol: $-OH$ attatched to CH, can react once and will react in both the oxidation and the lucas test
- 3° Alcohol: $-OH$ attatched to C, cannot react but will only react in the lucas test

4.12 Aldehydes

Aldehydes are formed by the oxidation of a 1° alcohol as follows



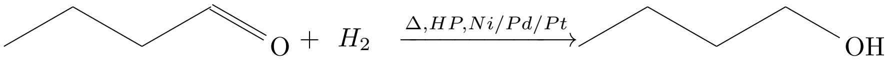
Since aldehydes stll has the oxygen, they can H-Bond with water and thus have high solubility. But among themselves, their only IMF is DD. Aldehydes are granted the highest numbering priority. Aldehydes are named with the "-al" suffix and a number should be added under no circumstances. For example



This has 4 carbons and is an aldehyde. Therefore it is called butanal.

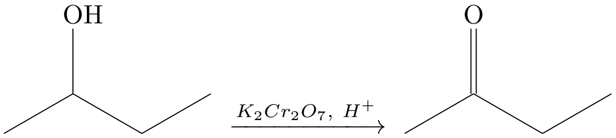
4.12.1 Reactions of Aldehydes

Aldehydes can be reduced to a 1st degree alcohol in a hydrogenation (reduction) reaction

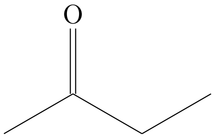


4.13 Ketones

Ketones are formed by the oxidation of a 2° alcohol as follows



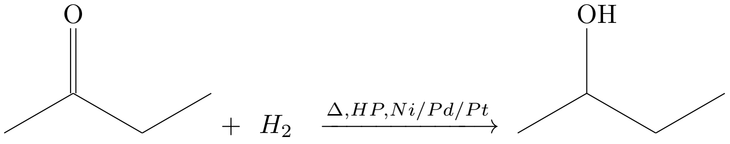
Since Ketones stll has the oxygen, they can H-Bond with water and thus have high solubility. But among themselves, their only IMF is DD. Ketones are granted the highest numbering priority. Ketones are named with the "-one" suffix and a number should be added to indicate the oxygen's location. For example



This has 4 carbons and has a =O at its 2nd carbon. Therefore it is called butan-2-one.

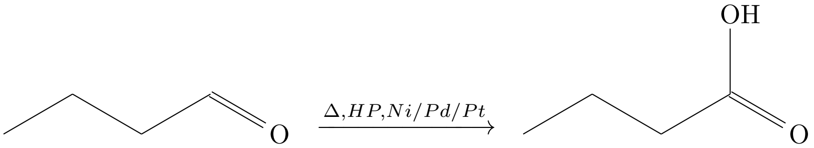
4.13.1 Reactions of Aldehydes

Ketones can be reduced to a 2nd degree alcohol in a hydrogenation (reduction) reaction

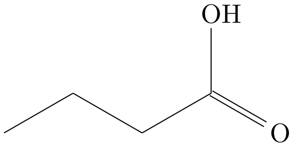


4.14 Carboxylic Acids

Carboxylic Acids are formed by the oxidation of a aldehyde, as represented by the following



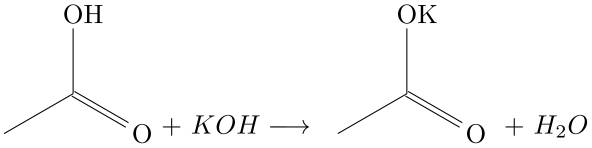
Since carboxylic acids stll has the oxygen, they can H-Bond with water and thus have high solubility. carboxylic acids are granted the highest numbering priority. carboxylic acids are named with the "-oic acid" suffix and a number should be added under no circumstances. For example



This has the carboxyl group and has 4 carbons. Therefore it is called butanoic acid

4.14.1 Reactions of Carboxylic Acids

Carboxylic acid can form salts with metals, as shown as follows



And these salts are named with the metal first and the original acid with the "-oate" suffix. For example, the above would be called potassium ethanoate.

4.15 Polymers

The only polymers studied in this course are addition polymers. Addition polymers are formed by addition reactions. And in an addition reaction, all the reactants must turn into the desired product and none are leftover. Polymers are made from individual monomers, and these monomers must have a C=C bond to open up to connect with other monomers to crate a polymer. Examples of polymers include PTFE, PVC, polyethene, polystyrene and polypropene

4.15.1 Crosslinking

Polymers can crosslink with other polymers to create larger polymers. POlymers with no crosslinks are held together by IMF dependent on side chains. Crosslinking creates more covalent bonds and thus makes the resultant polymer more rigid than before.

4.16 Common Names of Organic Compounds

- 1-methylbenzene: xylene
- 1,2-dimethyl benzene: o-xylene
- 1,3-dimethyl benzene: m-xylene
- 1,4-dimethyl benzene: p-xylene
- 2-methyl-1,3,5-trinitrobenzene: trinitrotoluene
- propan-2-one: acetone

- methanal: formaldehyde
- methanoic acid: formic acid
- ethanoic acid: acetic acid
- ethandioic acid: oxalic acid
- benzenol: phenol
- citric acid (IUPAC name not taught)
- propan-1-2-3-triol: glycerol
- ethanol: ethyl alcohol
- propan-2-ol: isopropanol

5 Unit 5

5.1 Oxidation and Reduction

When electron transfer is involved in a chemical reaction, it is called a oxidation and reduction (or redox) reaction. Redox reactions convert chemical potential energy into electrical energy through the movements of electrons. There are 2 parts to a redox reaction:

- Oxidation, the lost of electrons
- Reduction, the gain of electrons

Note that oxidising agents oxidise other things by taking their electrons (thus reducing itself) and reduction agents reduce other things by giving them electrons (thus oxidising itself). For example, consider the following reaction

$$Zn + I_2 \rightarrow ZnI_2$$

To find out the oxidising and reducting agents, the above equation must be written in its net ionic form

$$Zn + I_2 \rightarrow Zn^{2+} + 2I^-$$

Now it becomes clear that zinc lost electrons and iodine gained electrons. Thus zinc is oxidised and iodine is reduced. We also see that the electrons from zinc has gone to iodine. In a sense, zinc gave its electrons to iodine, and therefore zinc is the reducing agent and iodine is the oxidising agent. The neutral atom and the ion of the same element in a redox reaction is known as a redox couple. For example, the two redox couples in this reaction are Zn/Zn^{2+} and I/I^-.

5.1.1 Charge and Oxidation states

An oxidation state is the charge an atom in a covalent bond would have if the bond was broken and all the electrons went to the more electronegative atom, and are often denoted with the sign before the number. Some common oxidation states include

Atom	Common Oxidation State	Exceptions
H	+1	Mixed hydrides (state=-1)
O	-2	OF_2(state=+2) and H_2O_2(state=-1)
Halogens	-1	More electronegative halogens (state=+1)

This is only a list of states for the most common atoms. but for other atoms, their oxidation state will need to be solved for.

For other atoms, their oxidation state can be found by taking the charge of the compound that it is part of, then substituting in the above values for their corresponding atom, and solving for the unknown oxidation state. For example, the oxidation state for sulfur in H_2SO_4 can be easily solved as follows (Note that the oxidation state of atoms are denoted by S_n, where n is the atom)

$$\begin{aligned} 2S_H + S_S + 4S_O = 0 \\ S_S &= -4S_O - 2S_H \\ &= 0 - 4(-2) - 2(+1) = +6 \end{aligned}$$

And the oxidation state of sulfur S_S is determiend to be +6.

5.2 Balancing Reactions for Charge

Balancing reactions for charge consists of writing the reaaction in terms of its oxidation and reduction half reactions, and then multiplying the equations by a factor such that the electron terms cancel out when the equations are added back together. For example, consider the following reaction.

$$HNO_3 + H_3AsO_3 \to NO + H_3AsO_4 + H_2O$$

By solving for the oxidation states, it becomes clear that the hydrogen and oxygen are not part of the redox, and only the arsenic and nitrogen are. And writing the half reactions by oxidation states gives the following

$$N^{+5} + 3e^- \to N^{+2} \tag{1}$$

$$As^{+3} \to As^{+5} + 2e^- \tag{2}$$

So, to cancel out the electron terms, we simply need to multiply (1)by 2 and (2) by 3. Which gives the following

$$\begin{aligned} 2N^{+5} + 6e^- &\to 2N^{+2} \\ 3As^{+3} &\to 3As^{+5} + 6e^- \end{aligned}$$

Adding together the equations and simplifying, we find that the net ionic equation represents the actual equation.

$$\begin{aligned} 2N^{+5} + 6e^- + 3As^{+3} &\to 2N^{+2} + 3As^{+5} + 6e^- \\ 2N^{+5} + 3As^{+3} &\to 2N^{+2} + 3As^{+5} \end{aligned}$$

And so we take the coefficients and multiply it with its respective terms in the actual equation.

$$2HNO_3 + 3H_3AsO_3 \to 2NO + 3H_3AsO_4 + H_2O$$

Thus the equation has been balanced for charge. To balance for mass, we can simply multiply the H_2O term by our desired amount and scale the whole equation. When balancing for mass after charge do not alter the ratio of coefficents in the atoms that are involved in redox.

5.2.1 Assuming Acid Conditions

For some reactions, the mass is left unbalanced even after the charge is balanced, and there are no non-ionic terms (like the H_2O above) to adjust and balance for mass. For example

$$VO_2^+ + Zn \to VO^{2+} + Zn^{2+}$$

Fully balancing for charge gives

$$2VO_2^+ + 2Zn \to 2VO^{2+} + Zn^{2+}$$

Which is not balanced in terms of mass. For a situation like this, we can simply assume that the reaction occurs in an acid solution, and that H^+ and H_2O are present as follows. Generally, add H_2O to one side first and if needed, multiply the H_2O and add a required amount of H^+ to the other side (H_2O and H^+ cannot be on the same side).

$$2VO_2^+ + 2Zn + 4H^+ \to 2VO^{2+} + Zn^{2+} + 2H_2O$$

5.3 Reactivity

The reactivity of an atom is determined by its standard reduction potential (E^{θ}, measured in volts). A higher standard reduction potential means that an atom is more likely to be reduced. For example, lead has $E^{\theta} = -0.13V$ while silver has $E^{\theta} = 0.80$, and since silver has the higher potential, it is more likely to be reduced. And from that information, we can see that a reaction represented by the net ionic equation below is likely to happen

$$Pb + Ag^{2+} \rightarrow Pb^{2+} + Ag$$

On the contrary, And the reverse is very unlikely to happen since the lead needs to be reduced but has a lower reduction potential.

5.4 Voltaic Cells

Voltaic cells are a type of cell that uses redox reactions to convert chemical potential energy to electrical energy. They usually contain a cathode and an anode connected by a wire (the external circuit, usually with a voltmeter attatched) and a salt bridge (a "bridge" filled with a salt solution, usually potassium nitrate to prevent depolarisation). The cathode side contains the cathode electrode in a solution of the electrode's metal's ions (ex. if the electrode is made of Cu, then the solution will be full of Cu^{2+} ions, usually in the form of a nitrate solution, like $Cu(NO_3)_2$). The anode follows the same structure. Electrons typically flow from the anode solution into the anode electrode, then into the wire before travelling into the cathode electrode and into the cathode solution.

5.4.1 Distinguishing Between the Cathode and the Anode

The cathode is positively charged and houses the metal that is being reduced. The anode is negatively charged and houses the metal that is being oxidised. For example, a voltaic cell made with lead and silver will have silver being the cathode and lead being the anode, since silver has a higher E^{θ} and thus is more likely to be reduced. Note that some voltaic cells are uses gasses (like hydrogen) and need an inert electrode (like platinum) to work. These cells work no different than normal cells except the inert electrode is attached to the wire above and placed in the solution below.

5.4.2 Standard Electrode Potential

The standard electrode potential (SHE or E^{θ}_{cell}) is obtained by the following formula

$$E^{\theta}_{cell} = E^{\theta}_{cathode} - E^{\theta}_{anode}$$

For example, the silver-lead cell in the above example's SHE can be calculated as follows

$$\begin{aligned} E^{\theta}_{cell} &= E^{\theta}_{cathode} - E^{\theta}_{anode} \\ &= 0.80V - (-0.13V) \\ &= 0.93V \end{aligned}$$

Note that if E^{θ}_{cell} comes out negative, that means that the cathode and anode has been reversed.

5.4.3 Standard Cell Notation

A quick way to represent voltaic cells is through the standard cell notation. The format is as follows

Anode Atom / Anode Ion // Cathode Ion / Cathode Atom

For example, the silver-lead cell in the above examples can be represented by the following in cell notation

$$Pb \ / \ Pb^{2+} // \ Ag^{2+} \ / \ Ag$$

The left side represents the anode half cell and the right side represents the cathode half cell. If an inert electrode is used, simply write the inert electrode after a comma after the atom is mentioned.