


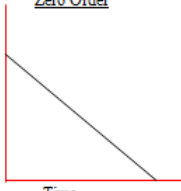
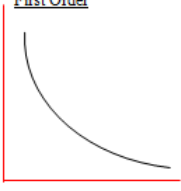
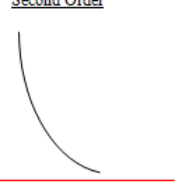


5.1 - Rates, equilibrium and pH

5.1.1 - How fast?

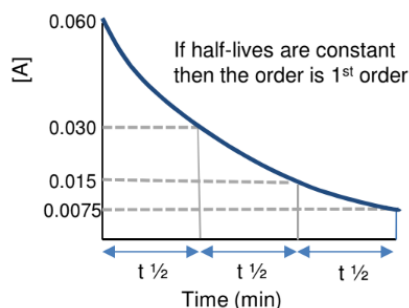
<p>What is rate of reaction?</p>	<p>The change in concentration of a substance per unit time.</p> <p><i>This is used for consistency as it's better than say change in mass per unit time.</i></p>
<p>Describe the following general rate equation: $r = k[A]^m[B]^n$</p>	<ul style="list-style-type: none"> Based on the reaction... $aA + bB \rightarrow \text{products}$. Unit of r (rate) is $\text{mol dm}^{-3} \text{s}^{-1}$, unit of $[A]$ and $[B]$ (conc. of A and B) are mol dm^{-3}, and unit of k (rate constant) varies. The rate constant is ONLY dependent on temperature (i.e., it increases with temperatures). The values m and n are called reaction orders and have nothing to do with the stoichiometric coefficients in the balanced equations. They are worked out experimentally and can be 0, 1, or 2. The overall order is $m + n$.
<p>What are the 3 orders of reaction with their associated rate-concentration and concentration-time graphs of the reactant(s)?</p>	<ul style="list-style-type: none"> Zero-order reaction - rate is independent of concentration. First-order reaction - rate \propto conc. Second-order rate - rate $\propto \text{conc}^x$ (by some factor). <div style="display: flex; flex-direction: column; align-items: center;"> <div style="display: flex; justify-content: space-around; width: 100%;"> <div style="text-align: center;"> <p><small>Rate vs. Time Graphs...</small></p> <p><u>Zero Order</u></p>  </div> <div style="text-align: center;"> <p><u>First Order</u></p>  </div> <div style="text-align: center;"> <p><u>Second Order</u></p>  </div> </div> <div style="display: flex; justify-content: space-around; width: 100%; margin-top: 10px;"> <div style="text-align: center;"> <p><small>Concentration vs. Time Graphs...</small></p> <p><u>Zero Order</u></p>  </div> <div style="text-align: center;"> <p><u>First Order</u></p>  </div> <div style="text-align: center;"> <p><u>Second Order</u></p>  </div> </div> <p><i>The concentration-time graph for a second-order reaction isn't required.</i></p> </div>
<p>What is the half-life of a reaction? And the half-life</p>	<ul style="list-style-type: none"> The time taken for the conc. of a reaction to half (which is constant for a first-order reaction).

equation for a first-order reaction?

- The equation for first order is:

$$t_{1/2} = \frac{\ln 2}{k}$$

- Shown on the graph:



*If the half-life **rapidly** increases then the reaction is second-order. If the half-life decreases then it's zero-order.*

How can you work out reaction orders using experimental initial rate data?

Experiment	Initial concentration of X/ mol dm ⁻³	Initial concentration of Y/ mol dm ⁻³	Initial rate/ mol dm ⁻³ s ⁻¹
1	0.05	0.1	0.15 x 10 ⁻⁶
2	0.10	0.1	0.30 x 10 ⁻⁶
3	0.20	0.2	2.40 x 10 ⁻⁶

- You have some equation $r = [X]^a[Y]^b$. Comparing (1) and (2), the order of reaction w.r.t. [X] is 1 (as initial rate doubles when initial conc. doubles).
- Comparing (2) and (3), both X and Y double whilst the initial rate increases x 8. Know [X] is order 1 so Y doubling increases initial rate x 4 \therefore order 2.

Say you have $r = k[A][B]^2$. The effect of individual changes in conc. are multiplied. E.g., if [A] x 2 and [B] x 3 then initial rate x 18.

Give 5 ways of determining the rate of reactions

- Use a gas syringe to measure the change in volume of gas.
- Use a top pan balance to measure the change in mass as gases escapes (works best with a heavy gas like CO₂).
- Measure the time taken for a precipitate to form.
- Measure a change in electricity conductivity (works well if ions are released).
- Use colorimetry to measure absorbance of a wavelength by a chemical.

What should be ensured when continually monitoring [A] and why? (in rates of reaction)	Having a large excess of everything else so their conc. is effectively constant \therefore its conc. won't change much during the reaction \therefore won't affect the reaction.
How can you use TLC to monitor the course of a reaction?	By taking samples at regular intervals.
Give 2 ways of determining orders of reaction and state the benefit of one over the other	<ol style="list-style-type: none"> 1. Continually monitor [A] and form a concentration-time graph for many different conc.'s then use this to plot a rate-concentration graph whose shape you can use. 2. Use a clock reaction to get a good approximation of the initial rate. <ul style="list-style-type: none"> • Then use the initial rates method. • The latter method is far less time consuming as a graph doesn't have to be plotted each time.
What 3 assumptions are made during a clock reaction?	<ol style="list-style-type: none"> 1. The concentration of reactant doesn't change significantly over the timescale. 2. Constant temperature. 3. When the endpoint is observed, the reaction has not proceeded too far. <p>This allows us to assume a constant rate.</p>
What is the rate determining step?	The slowest step in a reaction mechanism which determines the overall rate of reaction.
What things should you look out for when proposing mechanisms from overall reactions and rate equations?	<ol style="list-style-type: none"> 1. A catalyst appearing in the rate equation as it will reform later. 2. When the slowest step contains an intermediate (not a reactant). It should be replaced by substances that make up the previous step. <p>E.g.,</p>

	<div data-bbox="589 212 1206 594"> <p>Example 2 <u>overall reaction</u> $A + 2B + C \rightarrow D + E$</p> <p><u>Mechanism</u></p> <p>Step 1 $A + B \rightarrow X + D$ fast</p> <p>Step 2 $X + C \rightarrow Y$ slow</p> <p>Step 3 $Y + B \rightarrow E$ fast</p> <p>$r = k [X]^1 [C]^1$</p> <p>The intermediate X is not one of the reactants so must be replaced with the substances that make up the intermediate in a previous step</p> <p>$A + B \rightarrow X + D$</p> <p>$r = k[A]^1[B]^1[C]^1$</p> </div> <div data-bbox="667 604 1336 989"> <p>Example 4</p> <p>Using the rate equation $\text{rate} = k[\text{NO}]^2[\text{H}_2]$ and the overall equation $2\text{NO}(\text{g}) + 2\text{H}_2(\text{g}) \rightarrow \text{N}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$, the following three-step mechanism for the reaction was suggested. X and Y are intermediate species.</p> <p>Step 1 $\text{NO} + \text{NO} \rightarrow \text{X}$</p> <p>Step 2 $\text{X} + \text{H}_2 \rightarrow \text{Y}$</p> <p>Step 3 $\text{Y} + \text{H}_2 \rightarrow \text{N}_2 + 2\text{H}_2\text{O}$</p> <p>Which one of the three steps is the rate-determining step?</p> <p>Step 2 – as H_2 appears in rate equation and combination of step 1 and 2 is the ratio that appears in the rate equation.</p> </div> <p><i>You need both to be able to propose the mechanism.</i></p>
<p>What is the Arrhenius Plot, what can it be used to find, and what does it use?</p>	<div data-bbox="1117 1077 1268 1119"> $k = Ae^{-E_a/RT}$ </div> <div data-bbox="613 1161 971 1476"> </div> <div data-bbox="1089 1182 1320 1224"> $\ln k = \ln(Ae^{-E_a/RT})$ </div> <div data-bbox="1073 1276 1336 1318"> $\ln k = \ln A + \ln e^{-E_a/RT}$ </div> <div data-bbox="1073 1381 1287 1465"> $\ln k = \ln A + \frac{-E_a}{RT}$ </div> <div data-bbox="732 1497 1398 1619"> <p>Plot-Wise Approach { $\ln k = \left(\frac{-E_a}{R}\right) \cdot \left(\frac{1}{T}\right) + \ln A$ }</p> </div> <p>Used to determine the pre-exponential factor and activation energy using a set of temperatures and rate constants.</p>

5.1.2 - How far?

This builds upon the K_c content from 3.2.3.

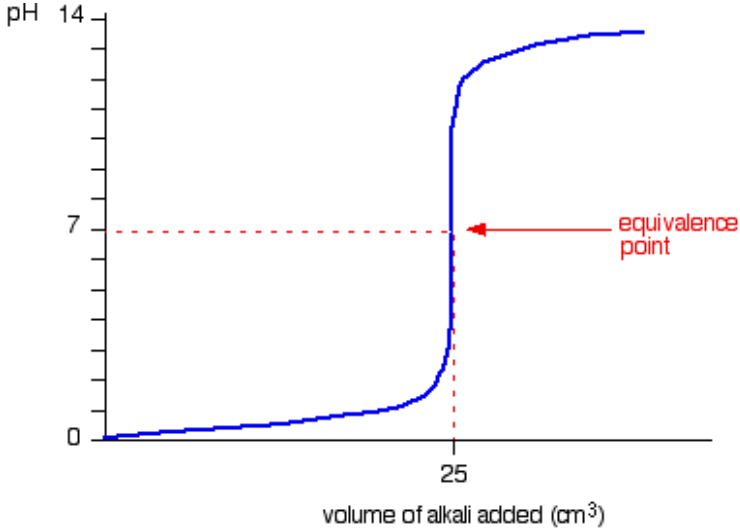
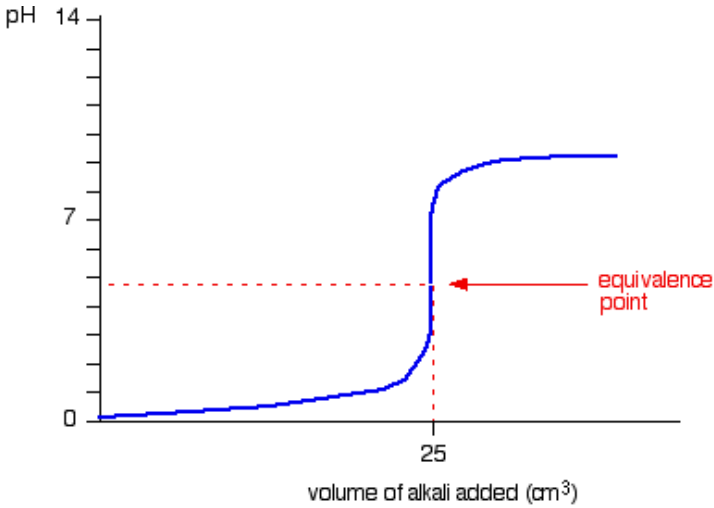
What method should you use for K_c questions?	<ul style="list-style-type: none"> Initial moles. Change in moles. Equilibrium moles.
What are the 2 types of equilibria?	<ul style="list-style-type: none"> Homogeneous - reactants and products are in the same physical states. Heterogeneous - reactants and products are in different physical states.
How do the K_c and K_p expressions change for heterogeneous equilibria?	<p>Only gases and aqueous reactants/products are used.</p> <p><i>E.g., $\text{Cu}_{(s)} + 2\text{Ag}^+_{(aq)} \rightleftharpoons \text{Cu}^{2+}_{(aq)} + 2\text{Ag}_{(s)}$ so $K_c = [\text{Cu}^{2+}] / [\text{Ag}^+]^2$.</i></p>
What are mole fractions and partial pressures? How do they relate?	<ul style="list-style-type: none"> Mole fraction of A - fraction of the total number of moles. Partial pressure of A - the pressure of A if it alone occupied the same volume vessel alone $p(\text{A}) = x_{\text{A}} \cdot P_{\text{tot}}$ <p><i>This means the total pressure is the sum of the partial pressures.</i></p>
What is K_p and why do we use it?	<ul style="list-style-type: none"> An equilibrium constant calculated from partial pressures. As it's difficult to get concentrations for gases. <p><i>E.g., $\text{H}_{2(g)} + \text{I}_{2(g)} \rightleftharpoons 2\text{HI}_{(g)}$ so $K_p = p(\text{HI})^2 / p(\text{H}_2)p(\text{I}_2)$.</i></p>
What method should you use for K_p questions?	ICE followed by $x_A P_A$.
What is the only thing that can affect K_c and K_p and why not the others?	<ul style="list-style-type: none"> Temperature. Changing the concentration/pressure of a substance won't have an effect as the system restores equilibrium so it's the SAME.

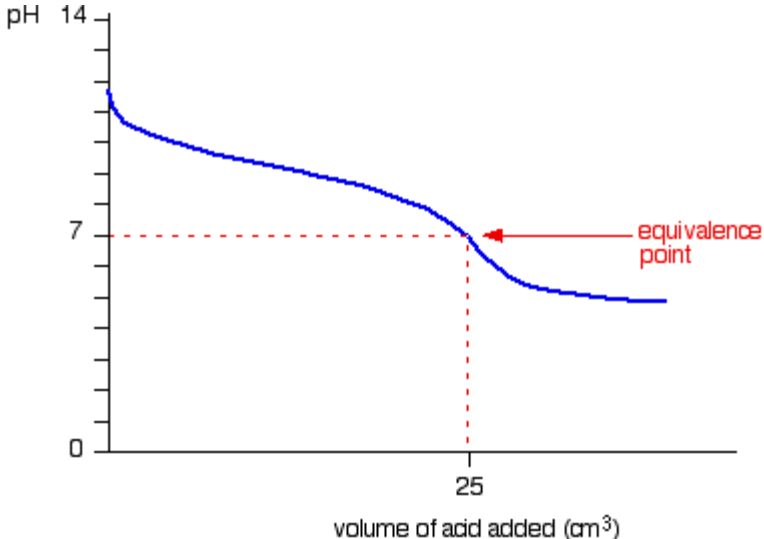
5.1.3 - Acids, bases and buffers

What is a Bronsted-Lowry acid/base? (use equations)	<ul style="list-style-type: none"> A substance that DONATES or ACCEPTS a proton. Base (B): $\text{B} + \text{H}_2\text{O} \rightleftharpoons \text{BH}^+ + \text{HO}^-$ Acid (A): $\text{H}_2\text{O} + \text{HA} \rightleftharpoons \text{H}_3\text{O}^+ + \text{A}^-$ <p>The H_3O^+ is called the hydronium ion but it is usually ignored.</p>
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<p>How do you find acid-base conjugate pairs?</p>	<p>Look forwards and backwards for substances that differ by H^+.</p> $HCOOH + CH_3(CH_2)_2COOH \rightleftharpoons HCOO^- + CH_3(CH_2)_2COOH_2^+$ <p style="text-align: center;">Acid 1 Base 2 Base 1 Acid 2</p> $HNO_3 + HNO_2 \rightleftharpoons NO_3^- + H_2NO_2^+$ <p style="text-align: center;">Acid 1 Base 2 Base 1 Acid 2</p> <p><i>The substances with a bigger K_a will act as an acid.</i></p>
<p>How do you calculate pH and pKa?</p>	$pH = -\log[H^+]$ $pK_a = -\log[K_a]$ <p><i>Using pKa allows for an easier comparison.</i></p>
<p>How does the pH and pKa of strong acids compare with weak?</p>	<p>The pH and pKa values for strong acids are lower as they fully dissociate (ie, $[H^+]_{final} = [HA]_{initial}$)</p>
<p>How can the weak acids be quantified? What assumptions are made in doing so?</p>	<p>Weak acids form an equilibrium mixture: $HA \rightleftharpoons H^+ + A^-$.</p> $K_a = \frac{[H^+][A^-]}{[HA]}$ <ol style="list-style-type: none"> 1. We can assume $[H^+] = [A^-]$ so all the H^+ ions come from the dissociation of the acid and not the water. 2. The dissociation is so small that $[HA]_{initial} = [HA]_{eqm}$ (this doesn't work for 'stronger' weak acids). <p><i>The first assumption doesn't work for buffers whilst the second doesn't work for 'stronger' weak acids.</i></p>
<p>How should you tackle a buffers question? (not a calculation)</p>	<ol style="list-style-type: none"> 1. Write down the equation: <ol style="list-style-type: none"> a. Eg, $HCOOH (high) \rightleftharpoons HCOO^- (low) + H^+ (low)$ 2. State what happens when acid is added: <ol style="list-style-type: none"> a. H^+ ions reacts with $HCOO^-$ so equilibrium (a) shifts left to make more $HCOOH$ (by Le Chatlier's). 3. State what happens when base is added: <ol style="list-style-type: none"> a. HO^- ions reacts with H^+ forms H_2O so equilibrium (a) shifts right to to replace H^+

<p>Derive the ionic product equation of water and when is it true</p>	<p>The following equilibrium is true for all (aq) substances and pure water: $\text{H}_2\text{O} (\text{l}) \rightleftharpoons \text{H}^+ (\text{aq}) + \text{HO}^- (\text{aq})$</p> $K_c = \frac{[\text{H}^+][\text{HO}^-]}{[\text{H}_2\text{O}]}$ $K_c \cdot [\text{H}_2\text{O}] = [\text{H}^+][\text{HO}^-]$ $K_w = [\text{H}^+][\text{HO}^-]$ <p>Since the $[\text{H}_2\text{O}] \gg [\text{H}^+]$, we can treat it as constant to form a new constant.</p>
<p>How can the ionic product of water be used?</p>	<ol style="list-style-type: none"> 1. To work out the pH of water since, for pure water, $[\text{H}^+] = [\text{HO}^-]$. 2. To convert between the $[\text{H}^+]$ or $[\text{HO}^-]$ for a dissolved substance. <p><i>Good for working out the $[\text{H}^+]$ ions in strong alkalis.</i></p>
<p>What MUST BE true for any neutral solution?</p>	<p>$[\text{H}^+] = [\text{HO}^-]$</p> <p>Since temp. changes the value of $K_w \Rightarrow [\text{H}^+]$ and $[\text{HO}^-]$ must change $\Rightarrow [\text{H}^+]$ is no longer $10^{-7} \Rightarrow \text{pH}$ no longer 7 $\Rightarrow \text{pH}$ of a neutral solution @ any temp. $\neq 7$. Ultimately, the pH scale changes with temp.</p>
<p>What does a buffer solution do?</p>	<p>Minimise pH changes (1) when <u>SMALL</u> amounts of acid or base is added (1)</p>
<p>Describe both ways of making acidic buffers using examples and comment on how this is similar for basic buffers</p>	<ol style="list-style-type: none"> 1. React a weak acid with a salt of its conjugate base: <ul style="list-style-type: none"> o Since $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$ and $\text{CH}_3\text{COO}^-\text{Na}^+ \rightarrow \text{CH}_3\text{COO}^- + \text{Na}^+$, combining the two forms $\text{CH}_3\text{COOH} \rightleftharpoons \text{CH}_3\text{COO}^- + \text{H}^+$ of which there is lots of weak acid and ions. 2. React an excess of weak acid with a strong base: <ul style="list-style-type: none"> o So you get $\text{CH}_3\text{COOH} + \text{NaOH} \rightarrow \text{CH}_3\text{COO}^-\text{Na}^+ + \text{H}_2\text{O}$ where the salt will dissociate. <ul style="list-style-type: none"> • Similarly, for basic buffers, you can react a weak base with a salt of its conjugate weak acid or react a weak base with a strong acid.

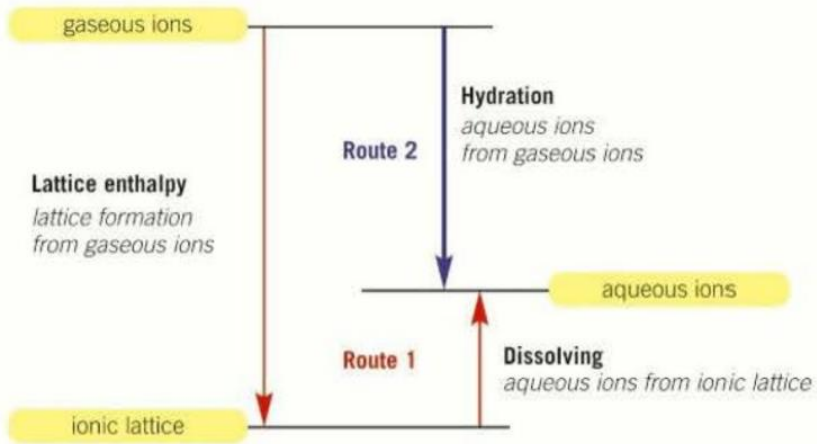
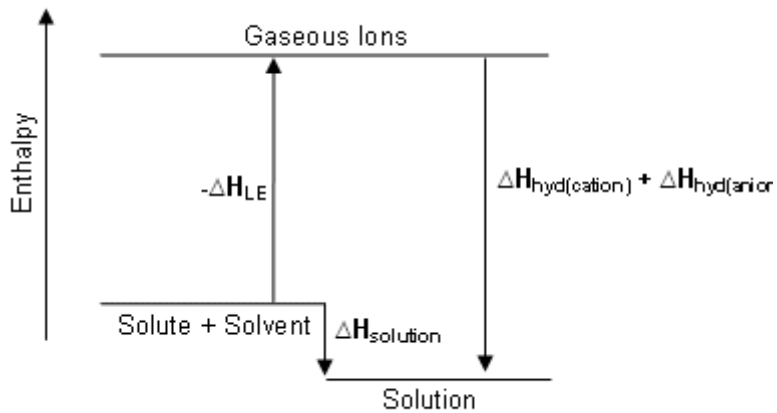
<p>How are buffers used in the body?</p>	<ul style="list-style-type: none"> • To keep the pH of the blood between 7.35 and 7.45. • An equilibrium exists between H_2CO_3 (carbonic acid) and HCO_3^- (hydrogen carbonate). • The amount of H_2CO_3 is controlled by respiration: $\text{H}_2\text{CO}_3 \rightleftharpoons \text{H}_2\text{O} + \text{CO}_2$
<p>What is the equivalence point of a titration?</p>	<p>When there are equal moles of acid and base.</p>
<p>Draw the titration curve for a strong-acid weak-base titration with the equivalence point and explain it using an example</p>	 <p>It is half the total volume in this case because of equal number of hydrogen and hydroxide ions from monoprotic acids and base.</p>
<p>Draw the titration curve for a strong-acid weak-base titration with the equivalence point and explain it using...</p> <p>$\text{HCl} + \text{NH}_4^+\text{HO}^- \rightarrow \text{NH}_4^+\text{Cl}^- + \text{H}_2\text{O}$</p>	 <p>$\text{HCl} + \text{NH}_4^+\text{HO}^- \rightarrow \text{NH}_4^+\text{Cl}^- + \text{H}_2\text{O}$ yet $\text{NH}_4^+\text{Cl}^- \rightarrow \text{NH}_4^+ + \text{Cl}^-$ so $\text{NH}_4^+ \rightleftharpoons \text{NH}_3 + \text{H}^+$ where the H^+ makes the solution slightly acidic.</p>

<p>Draw the graph of a weak-acid weak-base titration</p>	 <p>We call the equivalence point approximately 7 if they're both equally weak.</p>
<p>How should you choose an indicator for a titration appropriately?</p>	<p>Make sure the equivalence point lies within the colour change range.</p>
<p>Which indicator is required for a weak-acid weak-base titration?</p>	<p>None as they won't work as there's no sharp pH change.</p>
<p>How do indicators work?</p>	<ul style="list-style-type: none"> • Considered as weak acids. • $\text{HIn (aq)} \rightleftharpoons \text{In}^- \text{ (aq)} + \text{H}^+ \text{ (aq)}$ • Colour A \rightleftharpoons Colour B • Using Le Chatelier's, if H^+ acid is added, it will shift to the left and vice versa for HO^- base.
<p>Which is more acidic and why, CH_3COCl or CH_3COOH?</p>	<p>CH_3COCl as it will form HCl in water.</p>

5.2 - Energy

5.2.1 - Lattice enthalpy

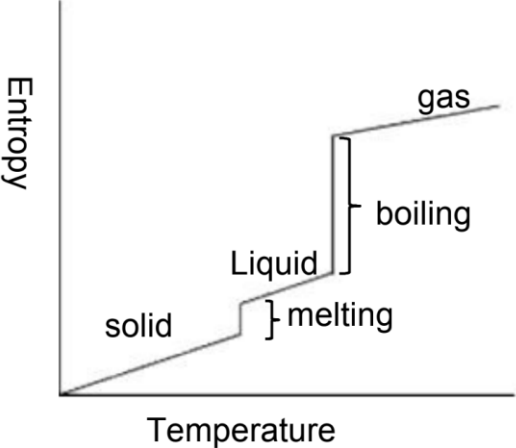
<p>What can lattice enthalpy be used as a measure of?</p>	<p>Ionic bond strength.</p>
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<p>What 2 factors affect lattice enthalpy and enthalpy change of hydration?</p>	<ul style="list-style-type: none"> • Charge difference between ions: <ul style="list-style-type: none"> ◦ Greater \Rightarrow stronger bonds \Rightarrow more energy released when broken \Rightarrow more exothermic enthalpy • Ionic radius: <ul style="list-style-type: none"> ◦ Smaller \Rightarrow stronger bonds \Rightarrow more energy released when broken \Rightarrow more exothermic enthalpy
<p>How is enthalpy of hydration related to enthalpy of solution?</p>	 <div style="border: 1px solid black; padding: 10px; margin-top: 10px;"> $\Delta_{\text{sol}}H = -\Delta_{\text{LE}}H + \sum \Delta_{\text{hyd}}H$ </div>
<p>When does a ionic substance generally dissolve?</p>	<p>If the hydration enthalpy > magnitude of lattice enthalpy so the overall enthalpy change is negative.</p>  <p>See above, the 'aqueous ions' will be lower than the 'ionic lattice'.</p>
<p>What does enthalpy change of solution tell you about solubility and why?</p>	<p>The more negative/exothermic the more soluble \because bond making is exothermic and bonds are made when water molecules bond to ions.</p>

<p>What are the key points / features of a Born-Haber cycle?</p>	<ul style="list-style-type: none"> • Downwards arrows - exothermic. Upwards arrows - endothermic. • May have to double or triple numbers occasionally. <p>Born Haber cycle: sodium Chloride</p> <p><i>Consider directions carefully - much like vector addition.</i></p>
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5.2.2 - Enthalpy and entropy

<p>What is entropy?</p>	<p>A measure of the dispersal of energy in a system OR how disordered something is.</p>
<p>What is a spontaneous process?</p>	<p>A process that proceeds on its own without external influence (so no continuous energy supply).</p> <p><i>Eg, diffusion.</i></p>
<p>Why can some endothermic reactions still be feasible?</p>	<p>As these substances become more energetically stable when there is disorder.</p>
<p>When is a substance thermodynamically stable?</p>	<p>When it reaches its maximum entropy state (lowest energy state).</p>

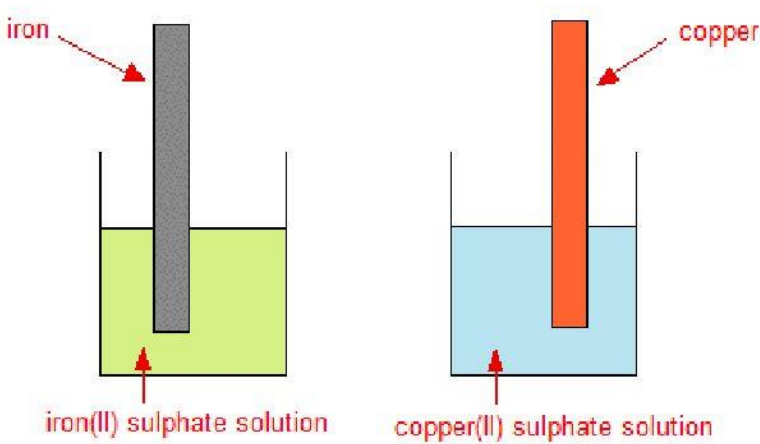
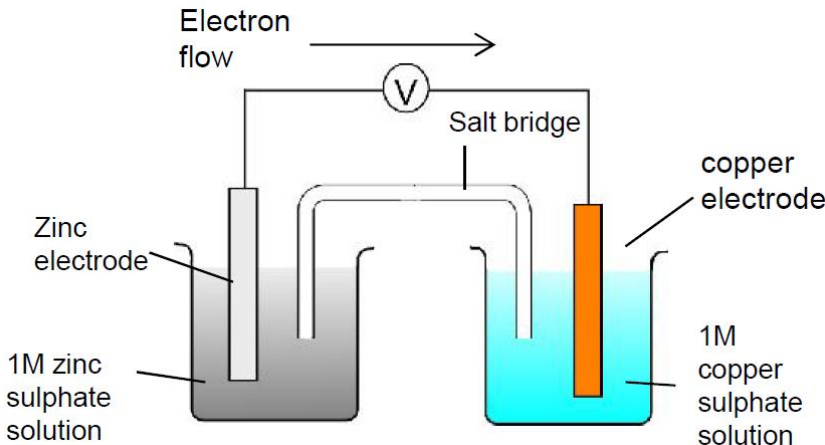
<p>Draw the temperature v. entropy graph</p>	
<p>How is entropy change (ΔS^\ominus) calculated? What is it measured in? And what does it being +ve and -ve mean?</p>	<ul style="list-style-type: none"> • $\Delta S^\ominus = \sum S^\ominus_{\text{products}} - \sum S^\ominus_{\text{reactants}}$ • Measured in $\text{JK}^{-1}\text{mol}^{-1}$ • +ve means increase in disorder/entropy. • -ve means decrease in disorder/entropy. <p><i>S^\ominus alone is the entropy per molecule which is always positive.</i></p>
<p>What is the Gibbs free energy equation? What are the outcomes?</p>	<ul style="list-style-type: none"> • $\Delta G = \Delta H - T\Delta S$ • $\Delta G < 0 \Rightarrow$ feasible (more negative \Rightarrow more feasible). • $\Delta G = 0 \Rightarrow$ equilibrium (this is similar for a puddle in equilibrium). • $\Delta G > 0 \Rightarrow$ not feasible. <p><i>Note that ΔS can be in $\text{JK}^{-1}\text{mol}^{-1}$ rather than $\text{kJK}^{-1}\text{mol}^{-1}$.</i></p>
<p>When is Gibbs free energy not useful?</p>	<p>While a reaction may be feasible, it can still...</p> <ul style="list-style-type: none"> • Have a very high activation energy. • Be incredibly slow that you won't it happening at all.
<p>How can Gibbs free energy be plotted? What does the gradient tell you?</p>	<ul style="list-style-type: none"> • Plot the ΔG v T of $\Delta G = \Delta H - T\Delta S$ as $y = mx + c$. • The gradient tells you the entropy change. <p><i>You can see when a reaction becomes feasible under this.</i></p>

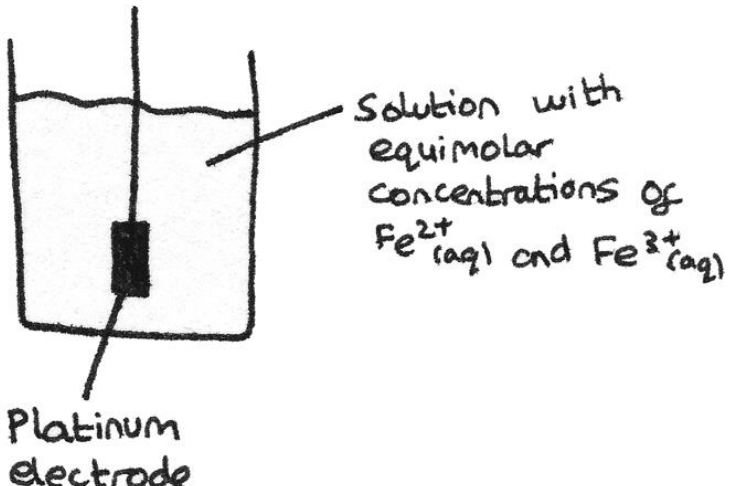
5.2.3 - Redox and electrode potentials

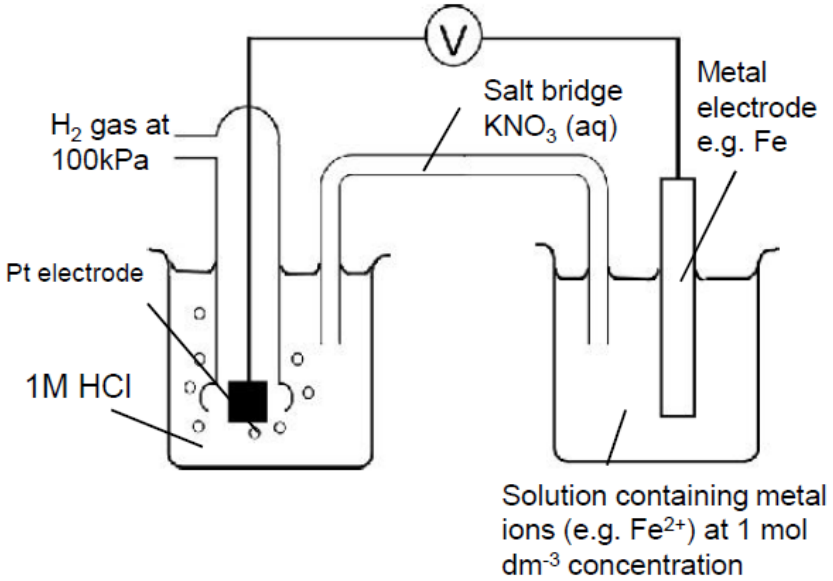
<p>How can <u>SIMPLE</u> half-equations be combined?</p>	<ol style="list-style-type: none"> 1. $\text{Br}_2 + 2\text{e}^- \rightarrow 2\text{Br}^-$ 2. $2\text{I}^- \rightarrow \text{I}_2 + 2\text{e}^-$
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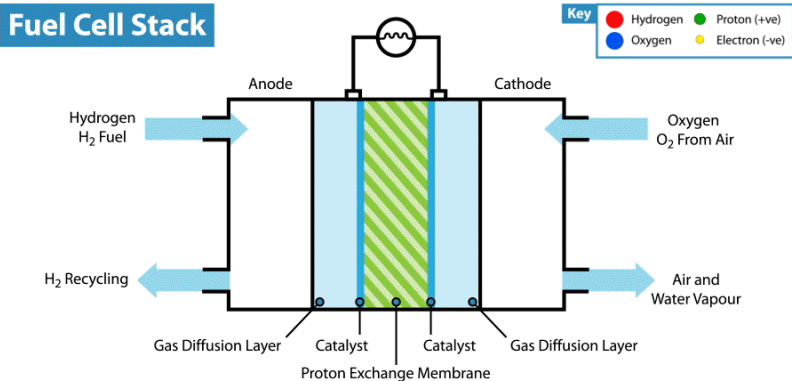
	<ul style="list-style-type: none"> These can be combined by adding them such that the electrons on either side cancel.
How can half-equations containing ^{12}O be balanced?	<ol style="list-style-type: none"> <u>O</u>xidation - balance oxidation numbers using e^-'s. <u>H</u>ydroxides and waters - balance atoms using HO^-'s and H_2O. <u>C</u>heck - check that atoms and charges balance. <p><i>This can be remembered as dodgy OHC.</i></p>
What is the manganate redox titration equation and its color change?	<ul style="list-style-type: none"> $\text{MnO}_4^- (\text{aq}) + 8\text{H}^+ (\text{aq}) + 5\text{Fe}^{2+} (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l}) + 5\text{Fe}^{3+} (\text{aq})$ From purple to colourless. As MnO_4^- is purple and Mn^{2+} is colourless.
What is different about the meniscus in a manganate redox titration?	The purple colour of the manganate can make it hard to see the bottom of the meniscus \Rightarrow measure from top of meniscus.
What acid should be used for manganate redox titrations? Why?	Dilute sulphuric acid \because other acid can set up alternative redox reactions (leading to inaccurate titres).
Why must excess sulphuric acid be used in a manganate redox titration?	<p>Otherwise, the solution won't be acid enough and MnO_2 will be produced instead of Mn^{2+}.</p> <p><i>Remember $\text{MnO}_4^- (\text{aq}) + 8\text{H}^+ (\text{aq}) + 5\text{Fe}^{2+} (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l}) + 5\text{Fe}^{3+} (\text{aq})$</i></p>
What is thiosulfate redox titration equation and its color change?	<ul style="list-style-type: none"> $2\text{S}_2\text{O}_3^{2-} (\text{aq}) + \text{I}_2 (\text{aq}) \rightarrow 2\text{I}^- (\text{aq}) + \text{S}_4\text{O}_6^{2-} (\text{aq})$ <ul style="list-style-type: none"> By combining $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2e^-$ and $\text{I}_2 + e^- \rightarrow 2\text{I}^-$. Left is yellow/brown solution. Right is colourless.
What is the problem with the endpoint of a thiosulfate redox titration? What should be done? Why does this work? (with colours)	<ul style="list-style-type: none"> The colour change is hard to spot. Starch indicator is added NEAR THE ENDPOINT to emphasise it (otherwise you get a false positive). It turns from from blue-black to colourless. Since with starch and I_2, you have blue-black. Yet once it is all used up (by adding thiosulfate), it becomes colourless. See below: <p>$2\text{S}_2\text{O}_3^{2-} (\text{aq}) + \text{I}_2 (\text{aq}) \rightarrow 2\text{I}^- (\text{aq}) + \text{S}_4\text{O}_6^{2-} (\text{aq})$</p>

How are thiosulfate reactions useful?	<p>You can determine the concentration of oxidising agent since $[O] + I^- \rightarrow I_2$ and then you can measure the conc of I_2 by doing a thiosulfate titration.</p> <p>Since $2S_2O_3^{2-} (aq) + I_2 (aq) \rightarrow 2I^- (aq) + S_4O_6^{2-} (aq)$.</p>
What will the stronger oxidising agent have in terms of E^0? Why?	<p>It will have a higher E^0 as it is a measure of how well a species gains e^-'s which an oxidising agent does as its reduced.</p>
What do standard electrode potentials tell us?	<ul style="list-style-type: none"> The measure of the ability of a species to gain electrons. The more positive E^0 is, the stronger the ability. When connected in an hydrogen electrode system. <div style="text-align: center;"> <p>oxidation</p> <p> $Li^+ + e^- \rightarrow Li$ $-3.03V$ $Mn^{2+} + 2e^- \rightarrow Mn$ $-1.19V$ $2H^+ + 2e^- \rightarrow H_2$ $0V$ $Ag^+ + e^- \rightarrow Ag$ $+0.8V$ $F_2 + 2e^- \rightarrow 2F^-$ $+2.87$ </p> <p>reduction</p> </div>
What do simple half-cells usually consist of?	<p>A metal (acting as an electrode) and a solution of compound containing said metal. Eg, Cu and $CuSO_4$.</p>

	 <p>iron</p> <p>iron(II) sulphate solution</p> <p>copper</p> <p>copper(II) sulphate solution</p>
<p>Why do half-cells produce a voltage?</p>	<p>The zinc half-cell has a tendency to oxidise / release e^-'s and copper half-cell has tendency to reduce / gain e^-'s leading to electrons flowing.</p>  <p>Electron flow</p> <p>V</p> <p>Salt bridge</p> <p>Zinc electrode</p> <p>1M zinc sulphate solution</p> <p>copper electrode</p> <p>1M copper sulphate solution</p> <ul style="list-style-type: none"> Thus... <ul style="list-style-type: none"> $Zn(s) \rightarrow Zn^{2+} + 2e^-$ $Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ <p><i>This is measured using a high-resistance voltmeter to stop current from flowing whilst preventing p.d. preventing reaction. If this was replaced with a bulb, current will flow. Reactions will then occur at each electrode yet the voltage would fall to zero as the reactants are used up.</i></p>
<p>When is a inert platinum electrode used in electrode</p>	<ul style="list-style-type: none"> When there is no metal surface for the reaction to take place on. This transfers e^-'s into or out of the cell. $Fe^{3+}(aq) + e^- \rightleftharpoons Fe^{2+}(aq)$

<p>potentials and why? (with example)</p>	
<p>Describe the salt bridge used in electrochemical cells (structure and reason)</p>	<ul style="list-style-type: none"> • Usually filter paper soaked in salt solution (eg, KNO_3). • Allows ions to form preventing imbalance (which would otherwise repel incoming metals). <p><i>The salt should be unreactive. Eg, KCl will form complexes instead.</i></p>
<p>Why isn't a wire used as the 'salt bridge'?</p>	<p>It would set up its own electrode system with solutions.</p>
<p>What is the standard hydrogen electrode, its equation and its conditions?</p>	<ul style="list-style-type: none"> • A half-cell using hydrogen. • Has the following equilibrium: $\text{H}_2 (\text{g}) \rightleftharpoons 2\text{H}^+ (\text{aq}) + 2\text{e}^-$ • Has the following conditions: <ul style="list-style-type: none"> ○ $\text{H}_2 (\text{g})$ @ 1 atm. ○ Solution containing H^+ ions @ 1 M (usually HCl). ○ Temperature @ 298K.

	
<p>What is the standard electrode potential of a half-cell (E^\ominus)?</p>	<p>E.m.f (of a half-cell) compared with a standard hydrogen half-cell under standard conditions</p> <p><i>This is by convention, we cannot measure an absolute potential for obvious reasons.</i></p>
<p>How may the Standard Hydrogen Electrode (SHE) be used indirectly and why?</p>	<p>By using a different standard electrode (which is calibrated against the SHE) since it's easier to use.</p>
<p>How are electrode potentials represented? (with sign)</p>	<ul style="list-style-type: none"> • $\text{Li}^+ (\text{aq}) \mid \text{L} (\text{s})$ <u>OR</u> $\text{Li}^+ (\text{aq}) + \text{e}^- \rightleftharpoons \text{Li} (\text{s})$ • Oxidation <u>ALWAYS</u> going right to left. • Negative \Rightarrow favours oxidation (<u>LOSS</u>) \Rightarrow negative electrode potential. <div style="text-align: center;"> $\text{Mg}^{2+} (\text{aq}) + 2\text{e}^- \rightarrow \text{Mg} (\text{s}) \quad E = -2.37\text{V}$ $\text{Cu}^{2+} (\text{aq}) + 2\text{e}^- \rightarrow \text{Cu} (\text{s}) \quad E = +0.34\text{V}$ </div>
<p>How do you work out the electrode potential of a cell?</p>	<ul style="list-style-type: none"> • $E_{\text{cell}} = E_{\text{red}} - E_{\text{ox}}$ • The more positive - less positive since more positive half-cell will always reduce (gain electrons).

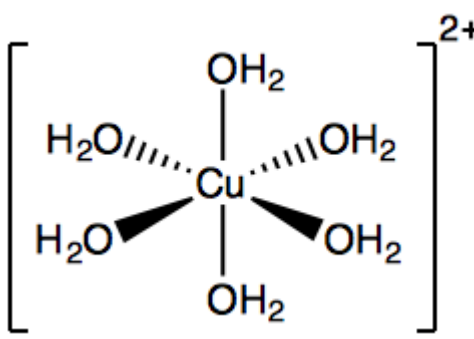
What does the emf of a cell represent in electrode potentials?	How far from equilibrium the cell lies (the more positive, the more likely the reaction is to occur).
How does concentration affect cell emf?	By Le Chatelier's, increasing [reactants] increases emf.
How does temperature affect cell emf?	Most cells are exothermic in the forwards direction. Using Le Chatelier's, increasing temperature decreases emf.
How are electrode potentials used to determine whether a reaction occurs? What's wrong with this?	<ul style="list-style-type: none"> If the E_{cell} is positive, the reaction may occur. <p>Yet...</p> <ul style="list-style-type: none"> The rate may be so slow that it appears not to be happening. The reaction has a high activation energy.
What is a fuel cell?	A cell using energy from a reaction of fuel with oxygen to create a voltage.
Draw and describe the hydrogen fuel cell	<p>With a semi-permeable membrane in the middle to keep the fuel and O_2 apart.</p>  <p>The diagram illustrates a Fuel Cell Stack. It consists of an Anode on the left and a Cathode on the right, separated by a central Proton Exchange Membrane. Both electrodes are coated with a Catalyst and a Gas Diffusion Layer. Hydrogen (H_2) fuel enters from the left, and Oxygen (O_2) from air enters from the right. A circuit with a resistor connects the two electrodes. A key indicates: Red circle for Hydrogen, Blue circle for Oxygen, Green circle for Proton (+ve), and Yellow circle for Electron (-ve). Arrows show H_2 recycling on the left and Air and Water Vapour exiting on the right.</p>
How do fuel cells maintain a constant voltage?	They are continuously fed fresh O_2 and H_2 thus maintaining concentration of reactants.
Give 3 advantages and 3 disadvantages of hydrogen fuel cells	<p>Advantages:</p> <ul style="list-style-type: none"> + Less pollution, CO_2 (only H_2O produced). + More efficient. + Large supply of water <p>Disadvantages:</p> <ul style="list-style-type: none"> - Need pressurised containers to store (risk explosion). - Have a limited lifetime (require regular replacement/disposal) - Uses toxic chemicals in their production.

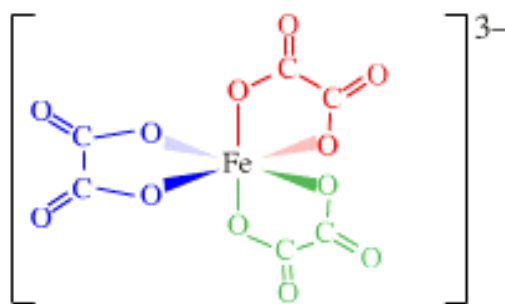
	<ul style="list-style-type: none"> - Hydrogen isn't readily available. - No H₂ infrastructure. 						
<p>What would happen if water was added to the Cu²⁺/Cu half-cell in the following redox system?</p> <table border="1"> <thead> <tr> <th>Redox system</th><th>E°/V</th></tr> </thead> <tbody> <tr> <td>Cu²⁺(aq) + 2e⁻ ⇌ Cu(s)</td><td>+0.34</td></tr> <tr> <td>Ag⁺(aq) + e⁻ ⇌ Ag(s)</td><td>+0.80</td></tr> </tbody> </table>	Redox system	E°/V	Cu ²⁺ (aq) + 2e ⁻ ⇌ Cu(s)	+0.34	Ag ⁺ (aq) + e ⁻ ⇌ Ag(s)	+0.80	<p>The cell potential increases (1) since...</p> <ul style="list-style-type: none"> • [Cu²⁺] decreases so equilibrium shifts left (1) • More electrons released by Cu (1) • Cell have bigger difference in E° (1). <p><i>This is from F325 June 2014 (7bii).</i></p>
Redox system	E°/V						
Cu ²⁺ (aq) + 2e ⁻ ⇌ Cu(s)	+0.34						
Ag ⁺ (aq) + e ⁻ ⇌ Ag(s)	+0.80						
What does each electrode do in fuel cells and why?	<ul style="list-style-type: none"> • Fuel always gives up its electrons (less +ve E°) at the negative anode (OXANA). • Oxygen always gains electrons (more +ve E°) at the positive cathode (REDCAT). 						

5.3 - Transition elements

5.3.1 - Transition elements

What is a transition element?	<p>A d-block element with at least one ion with an incomplete d-orbital.</p> <p>Eg, Cu has [Ar] 3d¹⁰4s¹ where Cu²⁺ has [Ar] 3d⁹4s⁰.</p>
Which transition metals have anomalous e⁻ structures and why?	<ul style="list-style-type: none"> • Copper ([Ar] 3d¹⁰4s¹) and chromium ([Ar] 3d⁵4s¹). • Having half-built/full-built 3d subshells is lower in energy than a full-built 4s subshell.
Why aren't zinc and scandium transition elements?	<ul style="list-style-type: none"> • Zinc can only form 2+ ion which has a complete d-subshell. <ul style="list-style-type: none"> ◦ Zn has [Ar] 3d¹⁰4s² whereas Zn²⁺ has [Ar] 3d¹⁰4s⁰. • Scandium can only form 3+ ions which has an empty d-subshell. <ul style="list-style-type: none"> ◦ Sc has [Ar] 3d¹4s² whereas Sc³⁺ has [Ar] 3d⁰4s⁰.
What are the characteristics of all transition elements?	<ul style="list-style-type: none"> • Formation of coloured ions. • Variable oxidation state. <ul style="list-style-type: none"> ◦ Eg, Fe (II) and Fe (III) since their 4s and 3d energy levels are so close, they can easily lose another e⁻. • Catalytic activity. • Complex ion formation in solution.

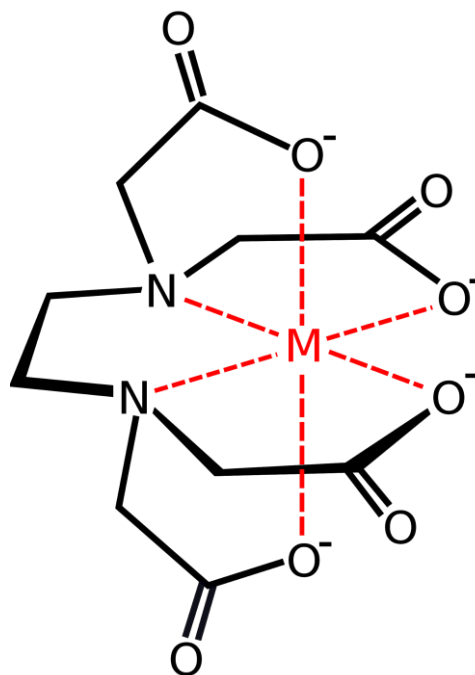
	<i>They form coloured ions due their electronic configuration.</i>
Why do transition elements make good catalysts?	They can change oxidation states by gaining/losing e ⁻ 's ⇒ can transfer e ⁻ 's to speed up reactions.
How are (aq) complex ions written?	<ul style="list-style-type: none"> When they contain ligands that aren't water or hydroxide ions, you need to include all water ligands. Eg, $[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ cannot be rewritten whereas $[\text{Cu}(\text{OH})_2(\text{H}_2\text{O})_4]^{2+}$ can be written as $\text{Cu}(\text{OH})_2$.
What is a complex ion?	A central metal ion surrounded by ligands.
What is a ligand?	An atom, ion, or molecule that can forms a coordinate bond to a central metal ion.
What is coordination number?	<p>The number of coordinate bonds formed to a central metal ions.</p> <p><i>This won't always equal the number of ligands.</i></p>
What is a monodentate ligand? (with examples of both ligands and complex ions)	<ul style="list-style-type: none"> Ligands that form 1 dative/coordinate bond. Examples include: H_2O, NH_3 and Cl^-. One is used as shown below: <div style="text-align: center;">  </div> <p><i>In this case, the water ligands have no charge so the charge stays as 2+ from Cu^{2+}.</i></p>
What is a bidentate ligand? (with examples of both ligands and complex ions)	<ul style="list-style-type: none"> Ligands that form 2 dative/coordinate bonds. Examples include: $\text{C}_2\text{O}_4^{2-}$ (ethanedioate), ethene-1,2-diamine / en. This is used as shown below with Fe (III).



Each ligand reduces the charge by 2. You have 3 hence 3- charge.

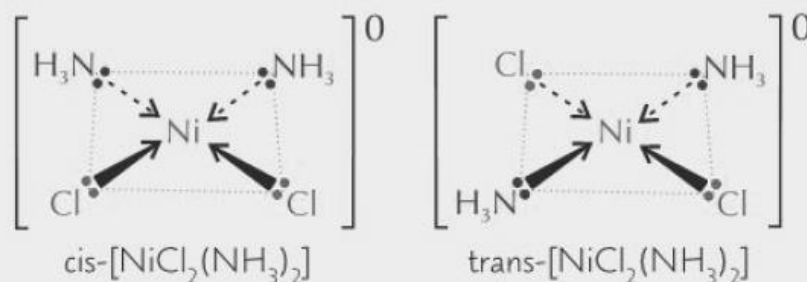
What is a multidentate ligand? (with an example)

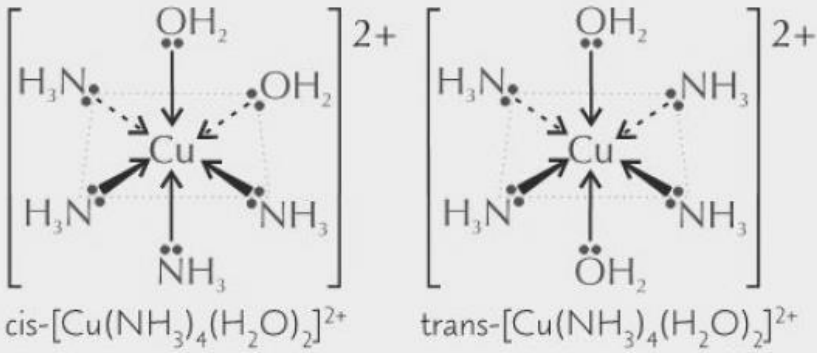
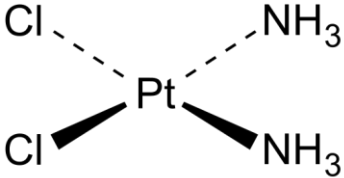
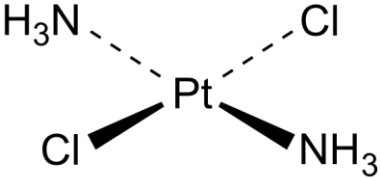
- Ligands that form several coordinate bonds.
- Examples include: EDTA^{4-} (forming 6 coordinate bonds).

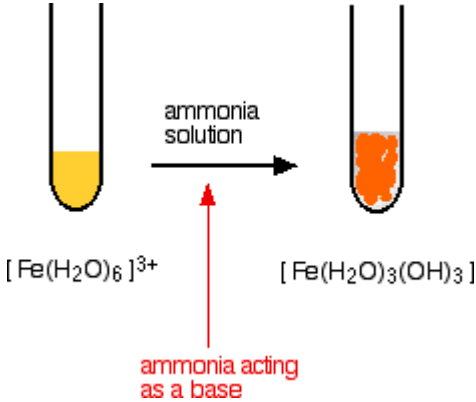


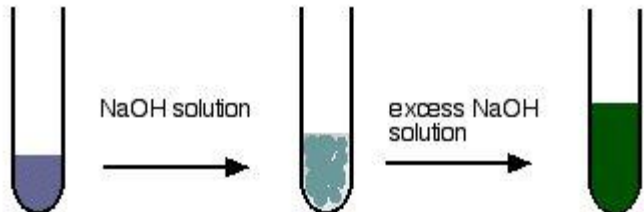
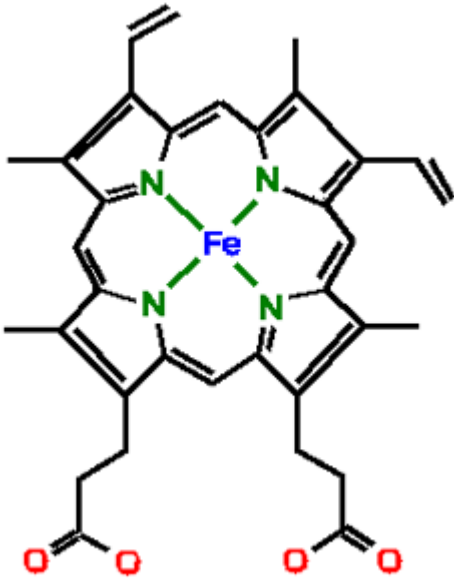
What complex ions show cis-trans isomerism?

Square planar and octahedral complex ions with at least two pairs of different ligands.



	 <p style="text-align: center;"> $\text{cis-}[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ $\text{trans-}[\text{Cu}(\text{NH}_3)_4(\text{H}_2\text{O})_2]^{2+}$ </p>
What is platin?	<p>A square planar molecule shown below.</p> <div style="display: flex; justify-content: space-around; align-items: center;"> <div style="text-align: center;">  <p>Cisplatin</p> </div> <div style="text-align: center;">  <p>Transplatin</p> </div> </div>
Why is cisplatin used as an anticancer drug? What are its drawbacks?	<ul style="list-style-type: none"> Prevents replication of cancer cells by binding to the DNA. Can prevent replication of healthy cells (\therefore suppresses immune system).
Describe the reaction of copper (II) with hydroxide ions (with colours)	<ul style="list-style-type: none"> $[\text{Cu}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 2\text{HO}^- (\text{aq}) \rightarrow [\text{Cu}(\text{H}_2\text{O})_4(\text{OH})_2] (\text{s}) + 2\text{H}_2\text{O} (\text{l})$ <ul style="list-style-type: none"> The complex loses its 2+ charge since HO^- has joined to it. Goes from pale blue to blue ppt (since no charge).
Describe the reaction of copper (II) with ammonia and excess ammonia (with colours)	<ul style="list-style-type: none"> $[\text{Cu}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 2\text{NH}_3 (\text{aq}) \rightarrow [\text{Cu}(\text{H}_2\text{O})_4(\text{OH})_2] (\text{s}) + 2\text{NH}_4^+ (\text{aq})$ <ul style="list-style-type: none"> The ammonia acts as a base accepting a H^+. Goes from pale blue to blue ppt. In excess NH_3, $[\text{Cu}(\text{H}_2\text{O})_2(\text{NH}_3)_4]^{2+} (\text{aq})$ is formed which is dark blue. <p><i>All 6 water molecules can be replaced with a very high conc. NH_3 but you don't need to know this.</i></p>

Which ligands are similar in size and which are different? What does this mean?	<ul style="list-style-type: none"> The ligands H_2O, NH_3 and CN^- are SIMILAR in size so coordination number doesn't change under substitution. The ligands H_2O and Cl^- are DIFFERENT (Cl^- is larger) in size so coordination number and shape changes under substitution.
Describe how copper (II) react with chloride ions (with colours)	<ul style="list-style-type: none"> $[\text{Cu}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 4\text{Cl}^- (\text{aq}) \rightleftharpoons [\text{CuCl}_4]^{2-} (\text{aq}) + 6\text{H}_2\text{O} (\text{l})$ <ul style="list-style-type: none"> Goes from blue to yellow. Reversible reaction so greenish colour
Describe how iron (II) react with hydroxide ions and ammonia (with colours)	<ul style="list-style-type: none"> $[\text{Fe}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 2\text{HO}^- (\text{aq}) \rightarrow [\text{Fe}(\text{OH})_2(\text{H}_2\text{O})_4] (\text{s}) + 2\text{H}_2\text{O} (\text{l})$ $[\text{Fe}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 2\text{NH}_3 (\text{aq}) \rightarrow [\text{Fe}(\text{OH})_2(\text{H}_2\text{O})_4] (\text{s}) + 2\text{NH}_4^+ (\text{aq})$ From pale green solution to green ppt.
Describe how iron (III) react with hydroxide ions and ammonia (with colours)	<ul style="list-style-type: none"> $[\text{Fe}(\text{H}_2\text{O})_6]^{3+} (\text{aq}) + 3\text{HO}^- (\text{aq}) \rightarrow [\text{Fe}(\text{OH})_3(\text{H}_2\text{O})_3] (\text{s}) + 3\text{H}_2\text{O} (\text{l})$ $[\text{Fe}(\text{H}_2\text{O})_6]^{3+} (\text{aq}) + 3\text{NH}_3 (\text{aq}) \rightarrow [\text{Fe}(\text{OH})_3(\text{H}_2\text{O})_3] (\text{s}) + 3\text{NH}_4^+ (\text{aq})$ From yellow solution to orange ppt. 
Describe how manganese (II) react with hydroxide ions and ammonia (with colours)	<ul style="list-style-type: none"> $[\text{Mn}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 2\text{HO}^- (\text{aq}) \rightarrow [\text{Mn}(\text{OH})_2(\text{H}_2\text{O})_4] (\text{s}) + 2\text{H}_2\text{O} (\text{l})$ $[\text{Mn}(\text{H}_2\text{O})_6]^{2+} (\text{aq}) + 2\text{NH}_3 (\text{aq}) \rightarrow [\text{Mn}(\text{OH})_2(\text{H}_2\text{O})_4] (\text{s}) + 2\text{NH}_4^+ (\text{aq})$ From pale pink solution to pink ppt.
Describe how chromium (III) reacts with hydroxide ions and excess hydroxide ions (with colours)	<p>WITHOUT EXCESS:</p> <ul style="list-style-type: none"> $[\text{Cr}(\text{H}_2\text{O})_6]^{3+} (\text{aq}) + 3\text{HO}^- (\text{aq}) \rightarrow [\text{Cr}(\text{OH})_3(\text{H}_2\text{O})_3] (\text{s}) + 3\text{H}_2\text{O} (\text{l})$ From green solution to grey-green ppt. <p>WITH EXCESS:</p> <ul style="list-style-type: none"> $[\text{Cr}(\text{H}_2\text{O})_6]^{3+} (\text{aq}) + 6\text{HO}^- (\text{aq}) \rightarrow [\text{Cr}(\text{OH})_6]^{3-} (\text{aq}) + 6\text{H}_2\text{O} (\text{l})$ From green solution to dark green solution.

	 <p style="text-align: center;"> $[\text{Cr}(\text{H}_2\text{O})_6]^{3+}$ $\xrightarrow{\text{NaOH solution}}$ $[\text{Cr}(\text{H}_2\text{O})_3(\text{OH})_3]$ $\xrightarrow{\text{excess NaOH solution}}$ $[\text{Cr}(\text{OH})_6]^{3-}$ </p>
What is 'haem'?	<p>An Fe (II) complex with a multidentate ligand.</p> 
What does haemoglobin do and how?	<p>Enables O₂ to be transported in the blood as O₂ bonds to the Fe²⁺ ions in haemoglobin (and is released when required).</p>
How does CO affect haemoglobin?	<p>It forms a coordinate bond with haemoglobin stronger than O₂ (more stable) preventing O₂ from attaching.</p>
Describe how iron (II) is oxidised (with colours)	<ul style="list-style-type: none"> • Fe²⁺ → Fe³⁺ using acidified manganate (VII) which is H⁺/MnO₄⁻. • The reaction goes from purple (MnO₄⁻) to colourless (Mn²⁺). • Yet, Fe²⁺ (aq) is green and Fe³⁺ (aq) is brown. <p><i>Oxygen in the air can bring this change also.</i></p> <p><i>The equation below is not required according to the specification, relevant information will be provided to work it out.</i></p>

	<ul style="list-style-type: none"> ● Overall Equation: <ul style="list-style-type: none"> ○ $\text{MnO}_4^- (\text{aq}) + 8\text{H}^+ (\text{aq}) + 5\text{Fe}^{2+} (\text{aq}) \rightarrow \text{Mn}^{2+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l}) + 5\text{Fe}^{3+} (\text{aq})$ ● Formed From: <ul style="list-style-type: none"> ○ $\text{MnO}_4^- (\text{aq}) + 8\text{H}^+ (\text{aq}) + 5\text{e}^- \rightarrow \text{Mn}^{2+} (\text{aq}) + 4\text{H}_2\text{O} (\text{l})$ ○ $\text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + \text{e}^-$ <p><i>Oxygen in the air can bring this change also.</i></p>
Describe how iron (III) is reduced (with colours)	<ul style="list-style-type: none"> ● $\text{Fe}^{3+} (\text{aq})$ is brown and $\text{Fe}^{2+} (\text{aq})$ is green. <p><i>The equation below is not required according to the specification, relevant information will be provided to work it out.</i></p> <ul style="list-style-type: none"> ● $\text{Fe}^{3+} \rightarrow \text{Fe}^{2+}$ using I^- (from KI) ● Equation: $2\text{Fe}^{3+} (\text{aq}) + 2\text{I}^- (\text{aq}) \rightarrow \text{I}_2 (\text{aq}) + 2\text{Fe}^{2+} (\text{aq})$
Describe how Cr^{3+} is oxidised to $\text{Cr}_2\text{O}_7^{2-}$ (with colours)	<p>Step One:</p> <ul style="list-style-type: none"> ● Cr^{3+} in $[\text{Cr}(\text{OH})_6]^{3+}$ is oxidised to CrO_4^{2-} by warming with H_2O_2 in alkaline conditions. ● This goes from dark green to yellow. <p>Step Two:</p> <ul style="list-style-type: none"> ● Add dilute sulfuric acid to the chromate (VI) solution to produce dichromate (VI) $\text{Cr}_2\text{O}_7^{2-}$. ● This goes from yellow to orange. <p><i>The equation below is not required according to the specification, relevant information will be provided to work it out.</i></p> <ul style="list-style-type: none"> ● Step 1 full equation: <ul style="list-style-type: none"> ○ $3\text{H}_2\text{O}_2 + 2\text{Cr}(\text{OH})_6^{3+} \rightarrow 2\text{HO}^- + 2\text{CrO}_4^{2-} + 8\text{H}_2\text{O}$ ● Step 1 half equations: <ul style="list-style-type: none"> ○ $\text{H}_2\text{O}_2 + 2\text{e}^- \rightarrow 2\text{HO}^-$ ○ $2\text{Cr}(\text{OH})_6^{3+} + 4\text{HO}^- \rightarrow 2\text{CrO}_4^{2-} + 8\text{H}_2\text{O} + 6\text{e}^-$ ● Step 2 full equation: <ul style="list-style-type: none"> ○ $2\text{CrO}_4^{2-} + 2\text{H}^+ \rightarrow \text{Cr}_2\text{O}_7^{2-} + \text{H}_2\text{O}$
Describe how $\text{Cr}_2\text{O}_7^{2-}$ is reduced to Cr^{3+} (with colours)	<ul style="list-style-type: none"> ● $\text{Cr}_2\text{O}_7^{2-}$ is reduced to Cr^{3+} by acidified zinc. ● This goes from orange to green.

	<p><i>The equation below is not required according to the specification, relevant information will be provided to work it out.</i></p> <ul style="list-style-type: none"> ● <u>Full equation:</u> <ul style="list-style-type: none"> ○ $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 3\text{Zn} \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} + 3\text{Zn}^{2+}$ ● <u>Half equations:</u> <ul style="list-style-type: none"> ○ $\text{Zn} \rightarrow \text{Zn}^{2+} + 2\text{e}^-$ ○ $\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}$
Describe how Cu^{2+} is reduced to Cu^+ (with colours)	<ul style="list-style-type: none"> ● Cu^{2+} is reduced to copper (I) iodide by iodide ions, I^-. ● From pale blue to off-white ppt. <p><i>The equation below is not required according to the specification, relevant information will be provided to work it out.</i></p> <ul style="list-style-type: none"> ● <u>Full equation:</u> <ul style="list-style-type: none"> ○ $2\text{Cu}^{2+}(\text{aq}) + 4\text{I}^-(\text{aq}) \rightarrow 2\text{CuI}(\text{s}) + \text{I}_2(\text{aq})$
Describe the disproportionation of Cu^+ (equation and why it happens)	<ul style="list-style-type: none"> ● $2\text{Cu}^+(\text{aq}) \rightarrow \text{Cu}(\text{s}) + \text{Cu}^{2+}(\text{aq})$ ● This spontaneous disproportion happens since Cu^+ is unstable.

5.3.2 - Qualitative analysis

Covered elsewhere.