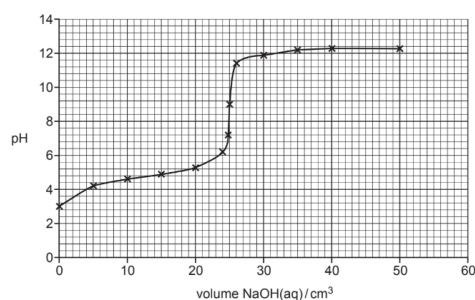


O.Q Exam questions from past papers

<p>Explain why many ionic substances are soluble in water, naming the bonds which are made and broken.</p> <p> In your answer you should use correct technical terms, spelled correctly.</p> <p>.....</p> <p>.....</p> <p>.....</p> <p>.....</p> <p>..... [4]</p> <p>Explain why many ionic substances are soluble in water,naming the bonds which are made and broken (4)</p>	<table><tr><td>electrostatic / ion-ion / ionic bonds broken ✓ ion-dipole bonds formed ✓ hydrogen bonds (in water) broken ✓ similar strength (between bonds broken and made) AW ✓ QWC both 'hydrogen bonds' and 'ion-dipole bonds' must be correctly spelled to score their marks</td><td>4 use of 'molecules' or 'atoms' to describe ionic substance CON: first mark IGNORE hydrogen bonds between ion and water IGNORE ionic dipole ALLOW bonds made are stronger than bonds broken' or 'energy released' AW ALLOW energy required to break bonds is less than energy required to make bonds' IGNORE 'intermolecular forces'</td></tr></table> <ul style="list-style-type: none">Ionic bonds and hydrogen bonds are broken (2)Ion-dipole bonds formed (1)Bonds made are stronger than bonds broken -> exothermic reaction (1) <p>Mark Schemes do not allow ionic dipole</p>	electrostatic / ion-ion / ionic bonds broken ✓ ion-dipole bonds formed ✓ hydrogen bonds (in water) broken ✓ similar strength (between bonds broken and made) AW ✓ QWC both 'hydrogen bonds' and 'ion-dipole bonds' must be correctly spelled to score their marks	4 use of 'molecules' or 'atoms' to describe ionic substance CON: first mark IGNORE hydrogen bonds between ion and water IGNORE ionic dipole ALLOW bonds made are stronger than bonds broken' or 'energy released' AW ALLOW energy required to break bonds is less than energy required to make bonds' IGNORE 'intermolecular forces'								
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<p>4 Some students investigate the dissolving of potassium salts.</p> <p>(a) The students are given 10.1 g of potassium nitrate, KNO₃. They dissolve this in 150 cm³ of water.</p> <p>The temperature goes down by 5.3 °C.</p> <p>(i) Give practical details of how they would carry out the experiment.</p> <p>Include the apparatus required and the measurements to be made.</p> <p>Describe the practical details of how a student would carry out to find the enthalpy change of solution of KNO₃(s) from having 150cm³ of water and 10.1g of KNO₃(s) (4)</p>	<ul style="list-style-type: none">Measure water using measuring cylinder (1)Use a polystyrene cup (1)Measure initial temperature (1)Then add solid and measure final temp when temperature change becomes constant (1) <table><tr><td>measure water using measuring cylinder ✓ insulated cup ✓ measure initial temp ✓ add solid and measure final temp <u>when all dissolved/when temperature (change) becomes constant</u> ✓</td><td>4 3.4 3.4 3.3 3.3 Eg polystyrene, polythene, styrofoam etc. (must be clear that this is before the addition of the solid)</td></tr></table>	measure water using measuring cylinder ✓ insulated cup ✓ measure initial temp ✓ add solid and measure final temp <u>when all dissolved/when temperature (change) becomes constant</u> ✓	4 3.4 3.4 3.3 3.3 Eg polystyrene, polythene, styrofoam etc. (must be clear that this is before the addition of the solid)								
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<p>Strontium oxide reacts with water to form strontium hydroxide, Sr(OH)₂.</p> <p>The solubility of strontium hydroxide in water at room temperature is around 10g dm⁻³.</p> <p>A student is given a saturated solution of strontium hydroxide, normal titration equipment and a variety of different concentrations of hydrochloric acid.</p> <p>The student wishes to find an accurate value for the concentration of the solution in mol dm⁻³.</p> <p>Describe in full a suitable procedure and indicate how the result would be calculated. [6]</p> <p>Describe in full a suitable procedure and indicate how the result would be calculated?</p>	<table><tr><td>Please refer to the marking instructions on page 4 of this mark scheme for guidance on how to mark this question. Level 3 (5 – 6 marks) Chooses an appropriate acid concentration.</td><td>6 3.4 x 3 3.3 x 3 Indicative scientific points include: Choice of acid concentration</td></tr></table> <hr/> <table><tr><th>Answer</th><th>Marks</th><th>AO element</th><th>Guidance</th></tr><tr><td>AND Gives a detailed description, including some fine detail, of procedure. AND Describes how the result would be calculated. <i>There is a well-developed line of reasoning which is clear and logically structured.</i> Level 2 (3 – 4 marks) Gives most of the key steps in the procedure, may include some fine detail AND describes how the result would be calculated. OR Addresses all three areas but lacks depth in any of them. <i>There is a line of reasoning presented with some structure. The information presented is relevant and supported by some evidence.</i> Level 1 (1 – 2 marks) A basic description of procedure. OR An attempt to describe the choice of acid concentration. 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<p>(iii) Explain how the solubility of strontium hydroxide in aqueous NaOH at 0 °C compares with $3.4 \times 10^{-2} \text{ mol dm}^{-3}$.</p> <p>Use the idea of solubility product in your answer.</p> <p>-----</p> <p>-----</p> <p>-----</p> <p>[2]</p> <p>Explain how the solubility of $\text{Sr}(\text{OH})_2$ in aqueous NaOH at 0 degrees compares with $3.4 \times 10^{-2} \text{ mol dm}^{-3}$ [Compared to dissolving in water]</p> <p>$K_{sp} = [\text{Na}^+][\text{OH}^-]$</p>	<table><tr><td>larger/increased concentration of OH^- ✓</td><td>2(AO3.2 x 2)</td><td>ALLOW more hydroxide ions</td></tr><tr><td>concentration of Sr^{2+} reduces in order for K_{sp} to remain constant AND solubility is lower ✓</td><td></td><td>ALLOW moves equilibrium to left AND solubility is lower Any reference to K_{sp} changing is CON</td></tr><tr><td></td><td></td><td>Examiner's Comments</td></tr><tr><td></td><td></td><td>This question proved to be a good question to differentiate between candidates. It needed the candidate to recognise that K_{sp} is a constant value and so whatever happens in the reaction mixture this has to remain unchanged.</td></tr><tr><td></td><td></td><td>Using the sodium hydroxide solution would result in a larger concentration of hydroxide ions being present, and so for K_{sp} to remain constant the concentration of the strontium ions had to reduce, which would result in the solubility of the strontium hydroxide decreasing.</td></tr><tr><td></td><td></td><td>A second acceptable route through this question was if a candidate recognised that the increase in hydroxide ions meant that the equilibrium established would act to counter the change by moving to the left, and that this would result in the solubility of the strontium hydroxide decreasing.</td></tr></table> <ul style="list-style-type: none">Larger concentration of OH^-Concentration of Sr^{2+} reduces in order for K_{sp} to remain constant and solubility is lower	larger/increased concentration of OH^- ✓	2(AO3.2 x 2)	ALLOW more hydroxide ions	concentration of Sr^{2+} reduces in order for K_{sp} to remain constant AND solubility is lower ✓		ALLOW moves equilibrium to left AND solubility is lower Any reference to K_{sp} changing is CON			Examiner's Comments			This question proved to be a good question to differentiate between candidates. It needed the candidate to recognise that K_{sp} is a constant value and so whatever happens in the reaction mixture this has to remain unchanged.			Using the sodium hydroxide solution would result in a larger concentration of hydroxide ions being present, and so for K_{sp} to remain constant the concentration of the strontium ions had to reduce, which would result in the solubility of the strontium hydroxide decreasing.			A second acceptable route through this question was if a candidate recognised that the increase in hydroxide ions meant that the equilibrium established would act to counter the change by moving to the left, and that this would result in the solubility of the strontium hydroxide decreasing.
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<p>$2\text{H}_2\text{O}_2(\text{l}) \rightarrow 2\text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$</p> <p>Equation 2.1</p> <p>The reaction in Equation 2.1 is accompanied by an increase in entropy.</p> <p>Give two reasons why an increase in entropy might have been expected.</p> <p>-----</p> <p>-----</p> <p>[2]</p> <p>Give two reasons why an increase in entropy might have been expected?</p>	<table><tr><td>On right (product side) there are more moles / molecules / particles ✓ gas on right / gases have more entropy / gas product AW✓</td><td>2</td><td>IGNORE 'more products'</td></tr><tr><td></td><td></td><td>Examiner's Comments</td></tr><tr><td></td><td></td><td>Candidates often scored the mark for a gas being produced but their terminology prevented scoring the second mark (more moles of products or more molecules formed), often writing just in terms of more products.</td></tr></table> <ul style="list-style-type: none">On right side there are more particlesA gas is on the right side which has more entropy than liquids	On right (product side) there are more moles / molecules / particles ✓ gas on right / gases have more entropy / gas product AW✓	2	IGNORE 'more products'			Examiner's Comments			Candidates often scored the mark for a gas being produced but their terminology prevented scoring the second mark (more moles of products or more molecules formed), often writing just in terms of more products.									
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(e) In a follow-up experiment, 25.0 cm³ of the ethanoic acid solution is titrated with a solution of sodium hydroxide of unknown concentration and the following graph is obtained.



Suggest a suitable practical procedure that would enable this graph to be obtained.

Use a

 [3]

Suggest a suitable procedure that would enable this graph to be obtained?

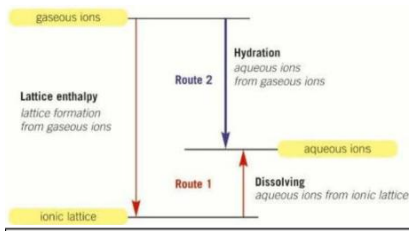
- Put NaOH in burette
- Pipette 25 cm³ of ethanoic acid into conical flask
- Measure pH with pH meter
- Run NaOH at 5 cm³ volumes at a time
- Run at 1 cm³ volumes when pH starts to change rapidly

- NaOH in burette
 - Pipette 25 cm³ ethanoic into beaker/conical flask
 - measure pH with a pH meter
 - run in 5 cm³ volumes at a time
 - smaller / 1 cm³ volumes when pH starts to change rapidly
- Any four scores three marks (✓✓✓)

3

O.a-h - Energetics | O1 | O4 | O5 |

What is a solute?	Something being dissolved into a solvent
What is a solvent?	The liquid in which a solute is dissolved in to form a solution
When does a solution USUALLY dissolve and not dissolve?	<ul style="list-style-type: none"> ● Dissolves when a reaction is exothermic where more energy is released when bonds are made ● Doesn't dissolve when a reaction is endothermic where more energy is required to break the bonds <p><i>Usually as its different if entropy applies</i></p>
When does an ionic substance generally dissolve?	<p>If the hydration enthalpy > lattice enthalpy so the overall enthalpy change is negative</p> <p><i>See above, the 'aqueous ions' will be lower than the 'ionic lattice'.</i></p>

	<p>exothermic lattice enthalpy</p> <ul style="list-style-type: none"> ● Ionic radius: Smaller ionic radius → attract water molecules more → stronger ion dipoles → more energy released → more exothermic lattice enthalpy
Why would LiF be more exothermic (more negative Δ_{LEH}) than NaF	<ul style="list-style-type: none"> ● Li compared to Na have the same charge ● Li ionic radius is smaller than Na and so would have a greater charge density → would attract the F more
Why would MgO be more exothermic (more negative Δ_{LEH}) than Na_2O ?	<ul style="list-style-type: none"> ● Mg compared to Na have a similar size ionic radius ● Mg ion has a 2+ charge compared to Na + charge and so would have a greater charge density → would attract the O more
What does enthalpy change of solution tell you about solubility and why?	The more negative/exothermic the more soluble a substance is → bond making is exothermic and bonds are made when water molecules bond to ions
How is enthalpy of hydration and lattice enthalpy related to enthalpy of solution?	 $\Delta_{\text{sol}}H = -\Delta_{LEH} + \sum \Delta_{\text{hyd}}H$ <p>For hydration enthalpies it's the sum of the hydration of the cation and anion</p>
Define Entropy (S)?	A measure of disorder of a chemical system and the number of ways of arranging the particles and their energy
How is the entropy change of a system (ΔS_{sys}) calculated? What is it measured in? And what does it being +ve and -ve mean?	<ul style="list-style-type: none"> ● $\Delta S_{\text{sys}} = \sum S_{\text{products}} - \sum S_{\text{reactants}}$ ● Measured in $\text{J K}^{-1} \text{mol}^{-1}$ ● +ve means increase in disorder/entropy ● -ve means decrease in disorder/entropy <p><i>S alone is the entropy per molecule which is always positive</i></p>

What is the trend in order of magnitude for the ΔS_{sys} for a solid, liquid or gas and why?	<ul style="list-style-type: none"> ● Gas > liquid > solid as gas particles has more arrangements and configurations than a liquid ● Gas particles have more energy than liquid particles and so would have more arrangements in filling up energy levels in terms of translational, rotational, vibrational or electronic levels <p><i>What is applied to comparing gas and liquids can be used to compare liquids and solids.</i></p> <p><i>Think of the analogy of giving people £1 compared to £5 to shop at Rio's, when people spend the money there is a higher chance the people with £1 would have the same money left compared to £5 as there are more spending configurations with £5 than £1 if you switch money with energy then you see - Mrs Marsden</i></p>
How is entropy change of surroundings (ΔS_{surr}) calculated?	<ul style="list-style-type: none"> ● $\Delta S_{\text{surr}} = -\Delta H/T$ ● Measured in $\text{J K}^{-1}\text{mol}^{-1}$ <p><i>Where T is in K. There is a negative sign as ΔH represents the system and since what you do to the surroundings will be opposite to the system there would be a - sign</i></p>
How is total entropy change (ΔS_{tot}) calculated? And what does it being +ve and -ve mean?	<ul style="list-style-type: none"> ● $\Delta S_{\text{tot}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}}$ ● +ve means the process is spontaneous ● -ve means the process isn't spontaneous ● Measured in $\text{J K}^{-1}\text{mol}^{-1}$ <p><i>Spontaneous means reactions that go of their own accord</i></p>
What is the significance of a reaction equilibrium mixture where $\Delta S_{\text{tot}} = 0$	<ul style="list-style-type: none"> ● The reaction is feasible ● The reaction is at equilibrium (therefore forward and backward reactions occur at the same rate) ● $K_c = 1$ <p><i>Feasible means reactions which are energetically favourable, a reaction that should go</i></p>
Define solubility product K_{sp} ?	<ul style="list-style-type: none"> ● Is when a sparingly soluble solid is dissolved in water to produce a saturated solution ● E.g $\text{Ba(OH)}_2 (\text{s}) \rightleftharpoons \text{Ba}^{2+} (\text{aq}) + 2\text{OH}^{-} (\text{aq})$ and therefore $K_{\text{sp}} = [\text{Ba}^{2+}_{\text{aq}}][\text{OH}^{-}_{\text{aq}}]^2$ <p><i>The maximum amount of solid which will dissolve in a certain solvent at a SPECIFIC TEMPERATURE</i></p>
What should you be aware of when tackling K_{sp} questions?	<ul style="list-style-type: none"> ● K_{sp} only changes with temperature and the [aq products] has to change to make it constant ● Putting state symbols in K_{sp} equation

	<ul style="list-style-type: none"> ● Comparing for e.g $\text{NaOH}_{(s)}$ and $\text{Be}(\text{OH})_2(s)$ where in $\text{NaOH}(s) \rightleftharpoons \text{Na}^+ + \text{OH}^-$ and $\text{Be}(\text{OH})_2 \rightleftharpoons \text{Be}^{2+} + 2\text{OH}^-$ so twice as many moles for OH^- ● Comparing for e.g $\text{NaOH}(s)$ and $\text{Be}(\text{OH})_2(s)$ where in K_{sp} for NaOH is $[\text{OH}]$ whilst in $\text{Be}(\text{OH})_2$ is $[\text{OH}]^2$
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O.i-m - Equilibria (acid-base) | O2 | O3 |

What is a Brønsted–Lowry acid/base? (use equations for strong acids/bases?)	<ul style="list-style-type: none"> ● A substance that DONATES or ACCEPTS a proton ● Base (B): $\text{B} + \text{H}_2\text{O} \rightarrow \text{BH}^+ + \text{OH}^-$ ● Acid (A): $\text{H}_2\text{O} + \text{HA} \rightarrow \text{H}_3\text{O}^+ + \text{A}^-$ ● Acid also: $\text{HA} \rightarrow \text{H}^+ + \text{A}^-$ <p>The H_3O^+ is called the oxonium ion/hydronium ion but it is usually ignored</p>
How do you find acid-base conjugate pairs?	<ul style="list-style-type: none"> ● Look forwards and backwards for substances that differ by H^+. ● And for substances that are the proton donors/proton acceptors $\text{HCOOH} + \text{CH}_3(\text{CH}_2)_2\text{COOH} \rightleftharpoons \text{HCOO}^- + \text{CH}_3(\text{CH}_2)_2\text{COOH}_2^+$ <p style="text-align: center;">Acid 1 Base 2 Base 1 Acid 2</p> $\text{HNO}_3 + \text{HNO}_2 \rightleftharpoons \text{NO}_3^- + \text{H}_2\text{NO}_2^+$ <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>
Define Conjugate acid and conjugate base?	<ul style="list-style-type: none"> ● Conjugate acid: Is a substance that can donate a proton to become a base (It was a base but accepted a proton to become a conj acid) ● Conjugate base: Is a substance that can accept proton to become an acid (It was an acid but donated a proton to become conj base) <p><i>E.g. $\text{HA} \rightleftharpoons \text{H}^+ + \text{A}^-$ If HA is the acid, A^- is the conj base and also HA is the conj acid of A^- and therefore A^- is the base in this case HA and A^- are the conjugate pairs</i></p>
Define strong acid?	A compound that fully dissociates in an acid to produce hydrogen ions [Is regarded as a proton donor]
What is the DEFINITION of pH/The equation for pH and pK_a of a strong acid?	$\text{pH} = -\log[\text{H}^+]$ $\text{pK}_a = -\log[\text{K}_a]$ <p><i>Using pK_a allows for an easier comparison</i></p> <p><i>When calculating H^+/K_a from pH/pK_a make sure to do $10^{-\text{pH}}/10^{-\text{pKa}}$ not $-10^{\text{pH}}/-10^{\text{pKa}}$</i></p>
Define strong base	A compound that fully dissociates when reacting with an

	acid to produce hydroxide ions [Is regarded as a proton donor]
What is the equation for the ionic product of water and when is it true?	$K_w = [H^+][HO^-]$ <p>The following equilibrium is true for all (aq) substances and pure water: $H_2O(l)$</p> <p><i>Since the $[H_2O]$ excess $\gg [H^+]$, we can treat it as constant to form a new constant</i></p>
What assumptions are there for strong bases when calculating pH	<p>All OH^- are from the base</p> <p><i>The OH^- from the dissociation of water is so low that its negligible</i></p>
How can the ionic product of water be used? $K_w = 1.00 \times 10^{-14} \text{ mol}^{-2} \text{ dm}^{-6}$ @ RTP	<ol style="list-style-type: none"> To work out the pH of water since, for pure water, $[H^+] = [OH^-]$ To convert between the $[H^+]$ or $[OH^-]$ for a dissolved substance. <p><i>Good for working out the $[H^+]$ ions in strong bases/alkalis</i></p>
What MUST BE true for any neutral solution?	<p>$[H^+] = [OH^-]$</p> <p>Since temp changes the value of $K_w \rightarrow [H^+]$ and $[OH^-]$ must change $\rightarrow [H^+]$ is no longer $10^{-7} \rightarrow$ pH no longer 7 \rightarrow pH of a neutral solution @ any temp. Not equal to 7. Ultimately, the pH scale changes with temp</p>
Define weak acid?	A compound that partially dissociates in water to produce hydrogen ions
What is the equation for weak acids? What assumptions are made in doing so?	<p>Weak acids form an equilibrium mixture: $HA \rightleftharpoons H^+ + A^-$.</p> $K_a = \frac{[H^+][A^-]}{[HA]}$ <ol style="list-style-type: none"> We can assume $[H^+] = [A^-]$ so all the H^+ ions come from the dissociation of the acid and not the water The dissociation is so small that $[HA]_{\text{initial}} = [HA]_{\text{eqm}}$ (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><i>K_a is known as the acidity constant or acid dissociation constant</i></p>

How does the pH and pK _a of strong acids compare with weak?	The pH and pK _a values for strong acids are lower as they fully dissociate (ie, $[H^+]_{final} = [HA]_{initial}$)
How would you calibrate a pH probe from textbook?	<ul style="list-style-type: none"> ● Wash the electrode with distilled water and then transfer the electrode into a buffer solution of pH = 7.00 ● To measure the pH of acidic solutions, calibrate the electrode using an acidic buffer solution typically with a pH = 4.00 ● To measure alkaline solutions calibrate the electrode using an alkaline buffer solution with pH = 10.00 ● Once calibrated you can measure the pH of any solution by rinsing the electrode then putting it in the solution to be measured
Define buffer and what does it do?	A substance that Resists changes in pH when <u>SMALL</u> quantities of acid or base are added
What are the assumptions when weak acids act as buffers?	<ol style="list-style-type: none"> 1. $[A^-] = [Salt]$ the concentration of the ion from the acid is the same as the salt 2. $[HA]_{eq} = [HA]_{initial}$ the equilibrium and initial concentration of acids are the same 3. A buffer doesn't stop pH from changing it only resists it so pH changes then is resisted
Describe the way of making acid buffers using an example?	<ol style="list-style-type: none"> 1. React a weak acid with its salt (salt would be conjugate base): E.g Ethanoic acid and sodium ethanoate ● The Ethanoic acid is a weak acid so $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ ● The salt fully dissolves and so $CH_3COO^-Na^+ \rightarrow CH_3COO^- + Na^+$ ● Combining the two forms $CH_3COOH \rightleftharpoons CH_3COO^- + H^+$ of which there is a lot of weak acid (ethanoic acid molecules) and ions (ethanoate ions from salt)
How should you tackle a buffer question? (not a calculation qu)	<ol style="list-style-type: none"> 1. Write down the definition of a buffer if it is not mentioned in the question 2. Write down the equation if not shown: <ol style="list-style-type: none"> a. Eg, $HCOOH$ (high conc) $\rightleftharpoons HCOO^-$ (low) + H^+ (low) 3. State what happens when acid is added: <ol style="list-style-type: none"> a. H^+ ions reacts with $HCOO^-$ POE shifts to the left to make more $HCOOH$ (by Le Chatliers). 4. State what happens when base is added: <ol style="list-style-type: none"> a. OH^- ions reacts with H^+ forms H_2O so POE shifts to the right to replace H^+ ions 5. State that pH remains constant as concentration of

	[HCOOH] and [HCOO ⁻] are both large
What is the expression for K _a for weak acids as BUFFERS and what does it mean?	<ul style="list-style-type: none"> ● $K_a = [H^+] \times \frac{[A^-]}{[HA]}$ ● [A⁻]/[HA] is a ratio so you don't always need the concentration for both substances also the ratio of concentrations for both does change even if the volume of salt differs for the same solution ● Altering [A⁻] or [HA] will affect how well a substance will buffer an acidic solution <p><i>Ratio [A⁻]/[HA] is usually between ¼ or 4/1 to buffer as any larger quantity would be more than a 'smaller quantity' so it wouldn't resist the change in pH</i></p>

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What is the effect on bonds when they absorb infrared radiation?	Bonds vibrate more
What is the main type of radiation that reaches the Earth from the sun (solar energy)? And how much does the earth absorb from it and what occurs after?	<ul style="list-style-type: none"> ● UV and Visible light ● Some of this energy is absorbed which heats up the earth ● The Earth radiates IR radiation
Name 3 greenhouse gases and what are their effects in the troposphere	<ul style="list-style-type: none"> ● Water vapour, methane and CO₂ ● They absorb some of the IR in the IR window <p><i>Not in the stratosphere haha and oxygen and ozone isn't a greenhouse gas</i></p>
What is the IR window?	<p>The wavelength of infrared radiation that water vapour does not absorb in the troposphere</p> <p><i>As water vapour is the most abundant greenhouse gas so the wavelengths it absorbs is the most important</i></p>
What happens when greenhouse gas molecules absorb extra IR radiation in the troposphere?	<p>The molecule's bond increases in vibrational energy to which the molecule transfers energy via colliding with other molecules which would increase their KE and raise the temperature</p> <p><i>Not in the stratosphere hahaha</i></p>
What occurs with some absorbed IR by the greenhouse gas molecules in the troposphere?	<ul style="list-style-type: none"> ● They re-emitted some of the absorbed IR in all directions some of which heats up the earth

	<i>It doesn't transmit or reflect it re-emits it and again not in the stratosphere hahahaha</i>
What occurs when the concentration of greenhouse gases increases?	This leads to an enhanced greenhouse effect