

ES.a-q - Elements from the sea | ES1-7

ES.Q Exam questions from past papers

<p>The Haber process $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \rightleftharpoons 2\text{NH}_3(\text{g})$ was formed due to Mrs Marsden and Mr Harbage giving their students impossible tests. The concentration for Ammonia = 32 mol dm^{-3} and $K_c = 2 \text{ mol dm}^{-3}$. What is the concentration of Nitrogen and Hydrogen?</p>	<p>Nitrogen - 2 mol dm^{-3} Hydrogen 2 mol dm^{-3}</p>
<p>Ammonia is not often made at temperatures below 473K. This is because the equilibrium is established too slowly at lower temperatures. Explain why the rate of a reaction increases with temperature. (2)</p> <p>(v) Ammonia is not often made at temperatures below 473K. This is because the equilibrium is established too slowly at lower temperatures.</p> <p>Explain why the rate of a reaction increases with temperature.</p> <p>.....</p> <p>.....</p> <p>.....</p> <p>..... [2]</p>	<p>Particles have more energy (1) and so there would be more frequent collisions with energy greater than the activation enthalpy (1)</p> <p>(c) (iv) Molecules/particles move faster/have more energy ✓ More (frequent) collisions with energy greater than activation enthalpy ✓</p> <p>2 1.2 'Atoms' CON first marking point ALLOW more successful collisions</p>
<p>Gallium has two isotopes, ^{69}Ga and ^{71}Ga. The Ar of gallium is 69.7. Calculate the relative abundance of ^{69}Ga as a percentage.</p> <p>(d) Gallium has two isotopes, ^{69}Ga and ^{71}Ga. The A_r of gallium is 69.7.</p> <p>Calculate the relative abundance of ^{69}Ga as a percentage.</p> <p>relative abundance of ^{69}Ga = % [2]</p>	<p>● $(69 \times x) + (71 \times (1-x)) = 69.7$ ● $69x + 71 - 71x = 69.7$ ● $2x = 1.3$ ● $x = 0.65$ ● $x = 65\%$</p> <p>22 d CHECK ANSWER ON ANSWER LINE 65% scores 2✓✓ $69x/100 + 71(100 - x)/100 = 69.7$ OR $69x + 71(1-x) = 69.7$ Answer = 65% ✓</p>
<p>(ii) All three products from the membrane cell are used in the chemical industry. Write the overall equation for the formation of these three products in the cell.</p> <p>$2\text{H}_2\text{O} + 2\text{NaCl} \rightarrow 2\text{NaOH} + \text{Cl}_2 + \text{H}_2$</p> <p>(iii) State and explain the atom economy of the reaction in (ii).</p> <p>State and explain the atom economy of the reaction? (1)</p>	<p>● 100% useful ● No waste ● All useful/desired products</p> <p>(iii) 100% since... no waste or all useful/desired products</p>

Commented [1]: Notice: For anyone using these flashcards feel free to learn definitions and equations but for other concepts (e.g. atom economy) the wording of the flashcard is suited to Destine needs, so please for the sake of your Grade in chemistry change the flashcard to suite your needs (Don't make life difficult on yourself)

Notice: Some of these flashcards may be difficult without context (especially those with mnemonics or anagrams) Come up to me (Destine) and ask!

Commented [2]: This is not fully correct needs to adjustments

$\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightleftharpoons 3\text{H}_2(\text{g}) + \text{CO}(\text{g}) \quad \Delta_r H = +206 \text{ kJ mol}^{-1}$ <p>Equation 33.1</p> <p>(e) The steam reforming reaction shown in equation 33.1 makes hydrogen. Much of the hydrogen is used in the manufacture of ammonia.</p> <p>(i) Calculate the atom economy of the reaction in equation 33.1 when making hydrogen.</p> <p>atom economy = % [1]</p> <p> $\text{CH}_4(\text{g}) + \text{H}_2\text{O}(\text{g}) \rightarrow 3\text{H}_2(\text{g}) + \text{CO}(\text{g}) \quad \Delta_r H = +206 \text{ kJ mol}^{-1}$ Equation 33.1 (e) The steam reforming reaction shown in equation 33.1 makes hydrogen. Much of the hydrogen is used in the manufacture of ammonia. (i) Calculate the atom economy of the reaction in equation 33.1 when making hydrogen. </p>	
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Question		Answer
(e)	(i)	$6 \times 100 / (16 + 18) = 17.6 / 17.65 / 18 \checkmark$
●	$(3 \times 2) / (12 + 4) + (2 \times 16) \times 100$	
●	$6 / (16 + 18) \times 100$	
●	17.65	

ES.a - Formulae, equations and amount of substance | ES6 |

What is atom economy and how is it calculated?	A measure of how efficient a reaction is (Relative formula mass not MASS in grams) $\text{atom economy} = \frac{\text{Mr of design products}}{\text{Mr of reactants}}$
What is a hydride?	A binary compound of hydrogen and a metal
What is a peroxide?	A compound containing two oxygen atoms bonded together (O-O)
What is an -ide?	A binary compound in which the nonmetal is given an -ide suffix (e.g., sodium oxide)
What do -ate compounds contain?	Oxygen and at least one other element <i>Examples include: Chlorate, nitrate and carbonate</i>
What is the charge of the ion Nitrite(III) and state oxidation numbers?	NO_2^-
What is the charge of the ion Sulfite(IV) and state oxidation numbers?	SO_3^{2-}
What is the charge of the ion Sulfate (VI) and state oxidation numbers??	SO_4^{2-}

What is the charge of the ion Nitrate (V) and state oxidation numbers??	NO_3^-
What is the charge of the ion Phosphate(V) and state oxidation numbers??	PO_4^{3-}
What is the charge of the ion Carbonate(IV) and state oxidation numbers??	CO_3^{2-}

ES.b-g - Redox | ES2 | ES3 | ES5 |

What is the oxidation number?	The no of electrons removed from an atom
What are the 4 definitions for oxidation?	<ul style="list-style-type: none"> ● The loss of electrons ● The gaining of oxygen atoms ● The increase in oxidation number ● The loss of hydrogen atoms (Learned about in WM with alcohols & phenols)
What are the 4 main rules for oxidation numbers?	<ol style="list-style-type: none"> 1. Elements always have an oxidation no. of 0 2. Sum of oxidation no. in a compound = 0 3. Sum of oxidation no. of a simple ion = its charge 4. Sum of oxidation no of a compound ion = its charge
What is a reduction reaction? What is a oxidation reaction?	<ul style="list-style-type: none"> ● Reduction reactions are the loss of oxygen atoms or gaining of electrons ● Opposite for oxidation <i>Remember OILRIG</i>
What are the 3 exceptions to oxidation rules?	<ol style="list-style-type: none"> 1. Hydrogen is always +1 except when in a hydride where its -1 2. Oxygen is always -2 except with peroxides where its -1 or F 3. Chlorine is always -1 except when bonded with oxygen or F
What is the mnemonic for remembering whether the anode and cathode is positive or negative?	Positive Anode Negative Is Cathode PANIC

How can <u>SIMPLE</u> half-equations be combined?	<ul style="list-style-type: none"> ● These can be combined by adding them such that the electrons on either side cancel ● $\text{Br}_2 + 2\text{e}^- = 2\text{Br}^-$ ● $2\text{I}^- = \text{I}_2 + 2\text{e}^-$
How can half-equations containing ^{12}O be balanced?	<ol style="list-style-type: none"> 1. Check that the element that is changing oxidation state is initially balanced i.e. $\text{MnO}_4^{2-} \rightarrow \text{Mn}^{2+}$ checking Mn atoms are balanced 2. Hydrogen ions and waters - balance atoms using H^+ and H_2O 3. Charges - balance charges using electrons 4. Check - check that atoms and charges balance. <p><i>This can be remembered as dodgy HOC</i></p> <p><i>You could also do the longer way by balancing an equation using OXIDATION STATES</i></p>
Which REDOX reaction occurs at the cathode and mnemonic?	<p>Reduction</p> <p><i>RED CAT</i></p>
Which REDOX reaction occurs at the anode and mnemonic?	<p>Oxidation</p> <p><i>OX ANA</i></p>
What products are made at a cathode of an electrolysis of solutions?	<ul style="list-style-type: none"> ● Hydrogen if the metal is above hydrogen in the reactivity series ● The Metal is produced for metals below hydrogen in the reactivity series ● Hydrogen also is made on electrolysis of acids
What products are made at an inert (unreactive) anode of an electrolysis of solutions?	<ul style="list-style-type: none"> ● Halogen if the salt is a halide ● Oxygen if the salt is a sulfate or nitrate ● Oxygen also is made on electrolysis of hydroxides
What is the metal e.g and half equation made at a reactive anode in $\text{CuSO}_4(\text{aq})$	<p>Copper and $\text{Cu}_{(\text{s})} = \text{Cu}^{2+}_{(\text{aq})} + 2\text{e}^-$</p>
What is the half equation at the inert anode if the salt is a sulfate or nitrate?	<p>$2\text{H}_2\text{O}(\text{l}) \rightarrow \text{O}_2(\text{g}) + 4\text{H}^+(\text{aq}) + 4\text{e}^-$</p>

What is the half equation at the inert anode if a hydroxide is being electrolysed?	$4\text{OH}^{-}(\text{aq}) \rightarrow \text{O}_2(\text{g}) + 2\text{H}_2\text{O}(\text{l}) + 4\text{e}^{-}$												
What is the half equation at the cathode if the metal in the salt is group 1 or 2 or is aluminium?	$2\text{H}_2\text{O}(\text{l}) + 2\text{e}^{-} \rightarrow 2\text{OH}^{-}(\text{aq}) + \text{H}_2(\text{g})$												
What is the half equation at the cathode on electrolysis of acids?	$2\text{H}^{+}(\text{aq}) + 2\text{e}^{-} \rightarrow \text{H}_2(\text{g})$												
What is an oxidising agent?	A substance that gains electrons and is reduced in a chemical reaction												
What is a reducing agent?	A substance that loses electrons and is oxidised in a chemical reaction												
Define Electrolysis	<p>A chemical change caused by passing an electric current through a compound which is either molten or in solution</p> <p><i>Don't need to remember but its good to be aware of in terms of chemical properties</i></p>												
Define Electrolyte	<p>A ionic compound in molten or solution that conducts electricity</p> <p><i>Don't need to remember but its good to be aware of in terms of chemical properties</i></p>												
What does starch solution test for?	Iodine												
What colour is an acid or an alkali if phenolphthalein is present?	<p>Colourless and pink</p> <table><thead><tr><th>Titration</th><th>Phenolphthalein</th><th>Methyl orange</th></tr></thead><tbody><tr><td>strong acid + weak base</td><td></td><td>Yes</td></tr><tr><td>weak acid + strong base</td><td>Yes</td><td></td></tr><tr><td>strong acid + strong base</td><td>Yes</td><td>Yes</td></tr></tbody></table>	Titration	Phenolphthalein	Methyl orange	strong acid + weak base		Yes	weak acid + strong base	Yes		strong acid + strong base	Yes	Yes
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strong acid + strong base	Yes	Yes											
What colour is an acid or an alkali if methyl orange is present?	<p>Red orange and yellow</p> <table><thead><tr><th>Titration</th><th>Phenolphthalein</th><th>Methyl orange</th></tr></thead><tbody><tr><td>strong acid + weak base</td><td></td><td>Yes</td></tr><tr><td>weak acid + strong base</td><td>Yes</td><td></td></tr><tr><td>strong acid + strong base</td><td>Yes</td><td>Yes</td></tr></tbody></table>	Titration	Phenolphthalein	Methyl orange	strong acid + weak base		Yes	weak acid + strong base	Yes		strong acid + strong base	Yes	Yes
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What is thiosulfate redox titration equation and its color change?	<ul style="list-style-type: none">● $2\text{S}_2\text{O}_3^{2-}(\text{aq}) + \text{I}_2(\text{aq}) \rightarrow 2\text{I}^{-}(\text{aq}) + \text{S}_4\text{O}_6^{2-}(\text{aq})$● By combining $2\text{S}_2\text{O}_3^{2-} \rightarrow \text{S}_4\text{O}_6^{2-} + 2\text{e}^{-}$ and $\text{I}_2 + \text{e}^{-} \rightarrow 2\text{I}^{-}$● Left is yellow/brown solution, right is colourless												

Commented [3]: Not on spec but good to know

Commented [4]: I really doubt this would come up

Commented [5]: Not on spec but good to know

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Commented [7]: I really doubt this would come up

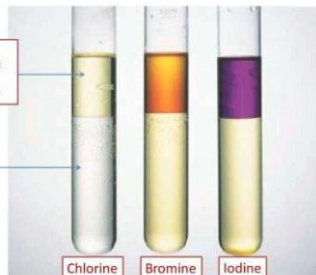
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What is the problem with the endpoint of a thiosulfate redox titration? What should be done? Why does this work? (with colours)	<ul style="list-style-type: none"> ● The colour change is hard to spot ● Starch indicator is added NEAR THE ENDPOINT to emphasise it (otherwise you get a false positive) ● It turns from 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: $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$
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 thiosulfate titration.</p> <p>Since $2S_2O_3^{2-}(aq) + I_2(aq) \rightarrow 2I^-(aq) + S_4O_6^{2-}(aq)$</p>
What should you look out for when doing oxidation/reduction questions (most likely 1 or 2 marks)	<p>For oxidation/reduction questions check if there are asking for electron transfer or oxidation number change and only include what the question requires</p> <p>I.e. $MnO_4^{2-} \rightarrow Mn^{2+}$</p> <p>Electron transfer: Mn loses 4 electrons</p> <p>Oxidation number change: Mn changes from +6 to +2</p>
What is the formula for percentage purity?	$\left(\frac{\text{Mass of sample}}{\text{mass of impure sample}} \right) \times 100$

ES.h-n - Inorganic chemistry and the periodic table | ES1 | ES5 | ES6 |

Give the colours and states of the first 4 halogens at Room Temperature? And what is the mnemonic to remember the colours	<p>F_2 - Pale yellow Gas Cl_2 - Green - Gas Br_2 - Red-brown - Liquid I_2 - Grey - Solid</p> <p>Remember: You Guys Read Gore. YELLOW GREEN RED GREY (down the group)</p>
What is the displacement of group 7 elements ?	More reactive halogen will displace a less reactive halide from its aqueous solution

	
<p>Give the test for halides with a precaution and example equation</p>	<ol style="list-style-type: none"> 1. Add ACIDIFIED silver nitrate solution 2. If the precipitate formed is white its silver chloride if its cream then its silver bromide and if its pale yellow then it is silver iodide  <p>Pot</p> <p>Silv</p> $Ag^+(aq) + Br^-(aq) \rightarrow AgBr(s)$ <p>To remember it learn the halides in alphabetical order BCI then the colours in alphabetical order CWY</p> <p>You could add instead nitric acid (to dissolve any carbonates aka a false positive which can mask desired observations) and silver nitrate solution</p>
<p>How can you be certain of the result of the halide test?</p>	<ol style="list-style-type: none"> 1. Add CONCENTRATED ammonia solution 2. Silver chloride precipitate dissolves 3. Silver bromide dissolves partially 4. Silver iodide does not dissolve at all  <p>Pot</p> <p>Silv</p> <p>Silver Iodide</p> <p>You must say CONCENTRATED as dilute is not allowed see below</p>

	<p>(ii) Describe what the student would do to the precipitates to distinguish between the halide ions in tubes A and B.</p> <p>State the expected results.</p> <p>.....</p> <p>.....</p> <p>.....</p> <p>(e) (ii) Add <u>concentrated</u> ammonia (solution) ✓</p> <table border="1"><tr><td></td><td>2</td><td>1.2 x 2</td><td>[2]</td></tr></table> <p>Bromide (partially) soluble AND iodide insoluble AW ✓</p> <p>MP2 depends on use of ammonia</p>		2	1.2 x 2	[2]
	2	1.2 x 2	[2]		
Why must experiments with halogens be done in a well-ventilated area?	They are powerful oxidising agents → toxic				
How will Br ₂ and I ₂ appear when in cyclohexane compared to water ?	<ul style="list-style-type: none">● Br₂ is orange in both● I₂ is orange/brown in water and violet in cyclohexane <div><p>Non-polar/hydrocarbon solvent example cyclohexane layer</p><p>Water layer</p><p>Chlorine Bromine Iodine</p><p>Fig. 2.5.13 Halogens are more soluble in organic solvents such as cyclohexane than in water.</p></div> <p>Br₂ being orange can be remembered by bromine water being orange</p>				

Commented [9]: Ray Amjad Example

What is the colour of chlorine when in water (aqueous) and in an organic solvent such as cyclohexane?

- Water (Aqueous) - Pale green
- Cyclohexane (Organic solvent) - Pale green

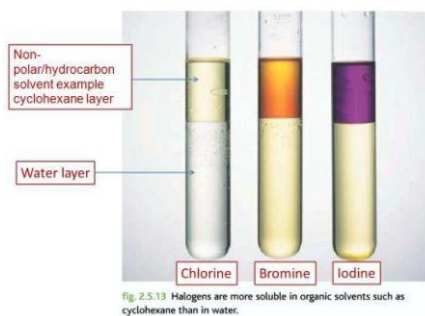
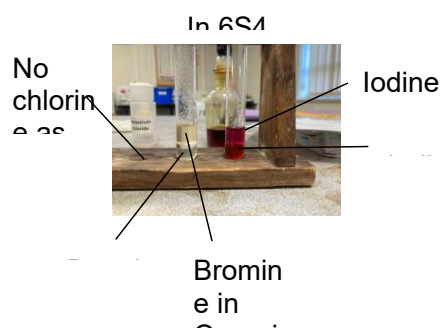


fig. 2.5.13 Halogens are more soluble in organic solvents such as cyclohexane than in water.

What is the colour of bromine when in water (aqueous) and in an organic solvent such as cyclohexane?

- Water (Aqueous) - Yellow
- Cyclohexane (Organic solvent) - Orange

No chlorine
Iodine
Bromine in

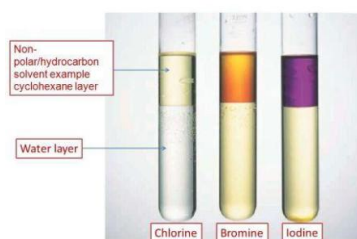
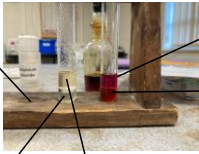
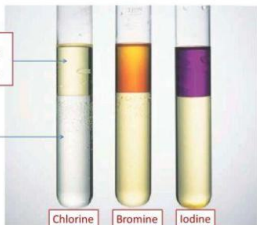
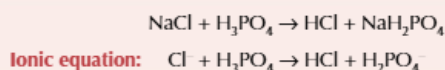


Fig. 2.5.13 Halogens are more soluble in organic solvents such as cyclohexane than in water.

<p>What is the colour of iodine when in water (aqueous) and in an organic solvent such as cyclohexane?</p>	<ul style="list-style-type: none"> ● Water (Aqueous) - Brown ● Cyclohexane (Organic solvent) - Purple <div data-bbox="619 521 991 763"> <p>No chlorine</p>  <p>Iodine</p> <p>Bromine</p> </div> <div data-bbox="611 786 960 1039">  <p>Non-polar/hydrocarbon solvent example cyclohexane layer</p> <p>Water layer</p> <p>Chlorine Bromine Iodine</p> <p><small>Fig 2.5.13 Halogens are more soluble in organic solvents such as cyclohexane than in water.</small></p> </div>
<p>What is the best and worst oxidising agent out of the halogens and why?</p>	<p>F₂ best and I₂ worst because its how easily a halogen can gain an electron</p>
<p>What is the best and worst reducing agent out of the halides and why?</p>	<p>I⁻ best and F⁻ worst because its how easily a halide can loss an electron</p>
<p>What is produced when Sodium halides (Chloride,Bromide,iodide) react with sulfuric acid and is it a redox reaction?</p>	<ul style="list-style-type: none"> ● White misty fumes (of HCl, HBr and HI) ● Not redox ● Its an impure product with NaBr and NaI and so H₂S, SO₂ or S are made <p>E.g $8\text{NaI} + \text{H}_2\text{SO}_4 \rightarrow \text{Na}_2\text{SO}_4 + 4\text{I}_2 + \text{H}_2\text{S} + 4\text{H}_2\text{O}$</p>

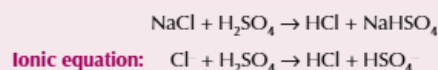
Hydrogen Halides can be made using Ionic Halides

- 1) You can make a hydrogen halide by adding a **concentrated acid**, to a solid, **ionic halide**
- 2) For example, to make **hydrogen chloride** (HCl), add **concentrated phosphoric acid** (H₃PO₄) to **sodium chloride**:

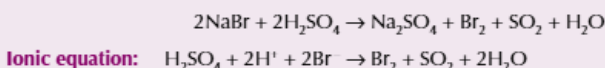


You don't need to learn the equations for the reaction of ionic halides with phosphoric acid — this is just here to show you what's going on.

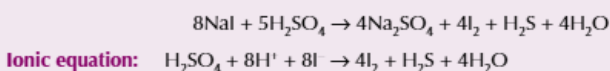
- 3) **All hydrogen halides** can be made this way, using an **ionic halide** and **concentrated phosphoric acid**.
- 4) Things get more complicated if you try to use **concentrated sulfuric acid** (H₂SO₄) to make hydrogen halides though. Unlike phosphoric acid, sulfuric acid is an **oxidising agent**, so it can get involved in redox reactions.
- 5) You **can** make **hydrogen chloride** using sulfuric acid:



- 6) You **can't** make either **hydrogen bromide** or **hydrogen iodide** using sulfuric acid.
- 7) When **sodium bromide** or **sodium iodide** react with **sulfuric acid**, the bromide or iodide ions are **oxidised** to make **bromine** or **iodine** gas.
- 8) This is because iodine and bromine are strong enough **reducing agents** to **reduce** sulfuric acid.
- 9) When you add **sulfuric acid** to **sodium bromide**, the **bromide ions** are oxidised to **bromine** gas and the **sulfuric acid** is reduced from **sulfuric acid** (oxidation state of S = +6) to **sulfur dioxide** (oxidation state of S = +4).



- 10) **Iodine** is such a strong reducing agent that when you add **sulfuric acid** to **sodium iodide** it reduces it all the way from **sulfuric acid** (oxidation state of S = +6) to **hydrogen sulfide** (oxidation state of S = -2).



Which Hydrogen halides do not react with sulfuric acid to make halides?

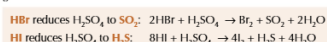
- HCl and HBr

What is the impurity balanced equation for if SO₂ forms from the reaction of Hydrogen Bromide with H₂SO₄? And what colour substance is formed?

- $2\text{HBr} + \text{H}_2\text{SO}_4 \rightarrow \text{Br}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$
- (Orange Vapour)
- Its both the ionic equation for when NaBr and H₂SO₄, and the equation when HBr reacts with H₂SO₄

Some Hydrogen Halides react with Sulfuric Acid

- 1) Hydrogen fluoride and hydrogen chloride **don't** react with **sulfuric acid** — they're not strong enough **reducing agents** to reduce the sulfur.
- 2) Hydrogen bromide and hydrogen iodide **do** react with sulfuric acid though.
- 3) The good news is that these reactions are similar to the sodium halide and sulfuric acid reactions on the last page:



<p>What is the impurity balanced equation for if SO₂ forms from the reaction of Hydrogen Iodide with H₂SO₄</p>	<ol style="list-style-type: none"> 1. $2\text{HI} + \text{H}_2\text{SO}_4 \rightarrow \text{I}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$ 2. Its both the ionic equation for when NaI and H₂SO₄, and the equation when HI reacts with H₂SO₄ <p>Some Hydrogen Halides react with Sulfuric Acid</p> <ol style="list-style-type: none"> 1) Hydrogen fluoride and hydrogen chloride don't react with sulfuric acid — they're not strong enough reducing agents to reduce the sulfur. 2) Hydrogen bromide and hydrogen iodide do react with sulfuric acid though. 3) The good news is that these reactions are similar to the sodium halide and sulfuric acid reactions on the last page: <p>HBr reduces H₂SO₄ to SO₂: $2\text{HBr} + \text{H}_2\text{SO}_4 \rightarrow \text{Br}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$ HI reduces H₂SO₄ to H₂S: $8\text{HI} + \text{H}_2\text{SO}_4 \rightarrow 4\text{I}_2 + \text{H}_2\text{S} + 4\text{H}_2\text{O}$</p>
<p>What is the impurity balanced equation for if S forms from the reaction of Hydrogen Iodide with H₂SO₄ and colours of substance produced</p>	<ul style="list-style-type: none"> ● $6\text{HI} + \text{H}_2\text{SO}_4 \rightarrow 3\text{I}_2 + \text{S} + 4\text{H}_2\text{O}$ ● (Yellow solid of S produced) ● Its both the ionic equation for when NaI and H₂SO₄, and the equation when HI reacts with H₂SO₄ <p>Some Hydrogen Halides react with Sulfuric Acid</p> <ol style="list-style-type: none"> 1) Hydrogen fluoride and hydrogen chloride don't react with sulfuric acid — they're not strong enough reducing agents to reduce the sulfur. 2) Hydrogen bromide and hydrogen iodide do react with sulfuric acid though. 3) The good news is that these reactions are similar to the sodium halide and sulfuric acid reactions on the last page: <p>HBr reduces H₂SO₄ to SO₂: $2\text{HBr} + \text{H}_2\text{SO}_4 \rightarrow \text{Br}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$ HI reduces H₂SO₄ to H₂S: $8\text{HI} + \text{H}_2\text{SO}_4 \rightarrow 4\text{I}_2 + \text{H}_2\text{S} + 4\text{H}_2\text{O}$</p>
<p>What is the impurity balanced equation for if H₂S forms from the reaction of Hydrogen Iodide with H₂SO₄</p>	<ul style="list-style-type: none"> ● $8\text{HI} + \text{H}_2\text{SO}_4 \rightarrow 4\text{I}_2 + \text{H}_2\text{S} + 4\text{H}_2\text{O}$ ● (Rotten egg smell) ● Its both the ionic equation for when NaI and H₂SO₄, and the equation when HI reacts with H₂SO₄ <p>Some Hydrogen Halides react with Sulfuric Acid</p> <ol style="list-style-type: none"> 1) Hydrogen fluoride and hydrogen chloride don't react with sulfuric acid — they're not strong enough reducing agents to reduce the sulfur. 2) Hydrogen bromide and hydrogen iodide do react with sulfuric acid though. 3) The good news is that these reactions are similar to the sodium halide and sulfuric acid reactions on the last page: <p>HBr reduces H₂SO₄ to SO₂: $2\text{HBr} + \text{H}_2\text{SO}_4 \rightarrow \text{Br}_2 + \text{SO}_2 + 2\text{H}_2\text{O}$ HI reduces H₂SO₄ to H₂S: $8\text{HI} + \text{H}_2\text{SO}_4 \rightarrow 4\text{I}_2 + \text{H}_2\text{S} + 4\text{H}_2\text{O}$</p>
<p>Which acid is used instead of Sulfuric acid to produce pure HBr and pure HI?</p>	<p>Phosphoric acid (H₃PO₄)</p> <p>Hydrogen Halides can be made using Ionic Halides</p> <ol style="list-style-type: none"> 1) You can make a hydrogen halide by adding a concentrated acid, to a solid, ionic halide 2) For example, to make hydrogen chloride (HCl), add concentrated phosphoric acid (H₃PO₄) to sodium chloride: <p>$\text{NaCl} + \text{H}_3\text{PO}_4 \rightarrow \text{HCl} + \text{NaH}_2\text{PO}_4$</p> <p>Ionic equation: $\text{Cl}^- + \text{H}_3\text{PO}_4 \rightarrow \text{HCl} + \text{H}_2\text{PO}_4^-$</p> <p><small>You don't need to learn the equations for the reaction of ionic halides with phosphoric acid — this is just here to show you what's going on.</small></p>
<p>What is the trend in thermal stability of the hydrogen halides as you go down group 7 and why</p>	<ul style="list-style-type: none"> ● Thermal stability decreases as bond length increases -> bond strength between the hydrogen and halide decreases -> less energy is needed to break the bond <p><i>Therefore HCl would have a higher average bond enthalpy and be more thermally stable compared to HI</i></p>

What is the equation for a Hydrogen Halide (Chloride) reacting with ammonia? (With state symbols?) and what colour fumes are produced	<p>$\text{HCl (g)} + \text{NH}_3 \text{ (g)} \rightarrow \text{NH}_4\text{Cl (s)}$ (White fumes)</p> <p>Hydrogen Halides react with Ammonia</p> <p>1) Ammonia (NH_3) is a base, so it can accept a proton to form the positively charged ammonium ion (NH_4^+) 2) The ammonium ion can bond with a negative halide ion, to produce an ammonium halide:</p> <div> $\text{HF}_{(\text{aq})} + \text{NH}_{3(\text{aq})} \rightarrow \text{NH}_4\text{F}_{(\text{aq})}$ $\text{HBr}_{(\text{aq})} + \text{NH}_{3(\text{aq})} \rightarrow \text{NH}_4\text{Br}_{(\text{aq})}$ $\text{HCl}_{(\text{aq})} + \text{NH}_{3(\text{aq})} \rightarrow \text{NH}_4\text{Cl}_{(\text{aq})}$ $\text{HI}_{(\text{aq})} + \text{NH}_{3(\text{aq})} \rightarrow \text{NH}_4\text{I}_{(\text{aq})}$ </div>
Explain the benefits and disadvantages of using chlorine to treat drinking water. Include references to the transport and storage of chlorine (Any 5 points)	<ol style="list-style-type: none"> 1. Kills bacteria/disinfectant 2. Cheap compared to other water treatment chemicals 3. Chlorine is a gas /making it difficult to contain/ so needs a strong container 4. Toxic/poisonous 5. Causes breathing problems 6. Forms by products 7. Dissolves in rivers /local water supplies, (dissolves in water) Forming bleach and acid (bleach and acid) kill life forms in the water
Give 3 large scale uses for the chlorine that is formed from brine/seawater?	<ol style="list-style-type: none"> 1. Bleach 2. Disinfectant /killing bacteria / sterilising 3. Extraction of bromine 4. Water treatment 5. Making PVC 6. Making solvents 7. Making HCl 8. Making medicines 9. Making pesticides 10. Make CFC's
Give one safety precaution that must be taken when working with chlorine gas?	<ol style="list-style-type: none"> 1. Well ventilated area 2. Fume cupboard 3. Breathing apparatus (say this as last resort as is not all mark schemes)
How would you test for a halogen ? (1)	<ul style="list-style-type: none"> ● Add organic solvent to see the halogens colour (1) ● Heat solution to see the colour of the halogens vapour (1) <p><i>NOT HALIDE so no silver nitrate solution easy mistake</i></p>

ES.o-q - Equilibria | ES4 | ES7 |

What is a closed system?	A container in which nothing can enter or leave except for heat
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What is dynamic equilibrium?	<ul style="list-style-type: none"> ● The rate of the forward reaction is equal to the rate of the backward reaction ● The concentration of reactants/products remain constant in a closed system
What is Le Chatelier's Principle?	If a dynamic equilibrium is changed the position of the equilibrium will oppose this change
What happens to the equilibrium position if the temperature is increased and why?	<ul style="list-style-type: none"> ● Favours the endothermic reaction ● To minimise the effect of temperature rise by absorbing heat
What happens to the equilibrium position if pressure is increased?	<ul style="list-style-type: none"> ● Favours the side producing the least moles of a gas ● To minimise the effect of pressure increase by reducing the number of particles <p><i>Remember that each mole takes 24dm³ under standard conditions!</i></p>
What is the equation for the equilibrium constant (K_c)?	$aA + bB \rightleftharpoons cC + dD$ $K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$
What does it mean if the value for K_c is less than 1?	<p>More reactants than products and so position of equilibrium shifts to the left</p> <p><i>This means that at equilibrium most of the products have been converted to reactant but not all</i></p>
What does it mean if the value for K_c is greater than 1?	<p>More products than reactants and so position of equilibrium shifts to the right</p> <p><i>This means that at equilibrium most of the reactants have been converted to products but not all</i></p>