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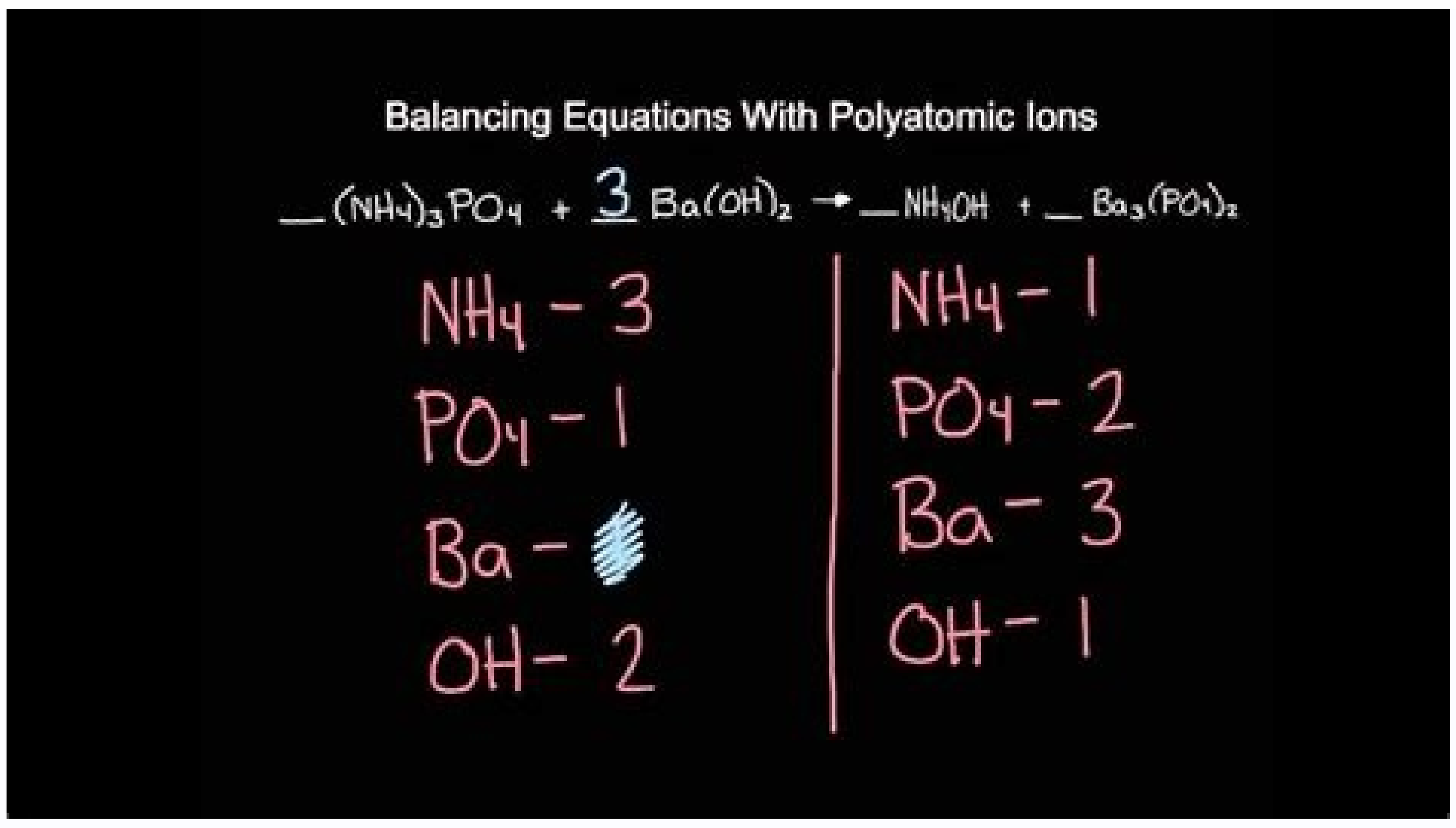

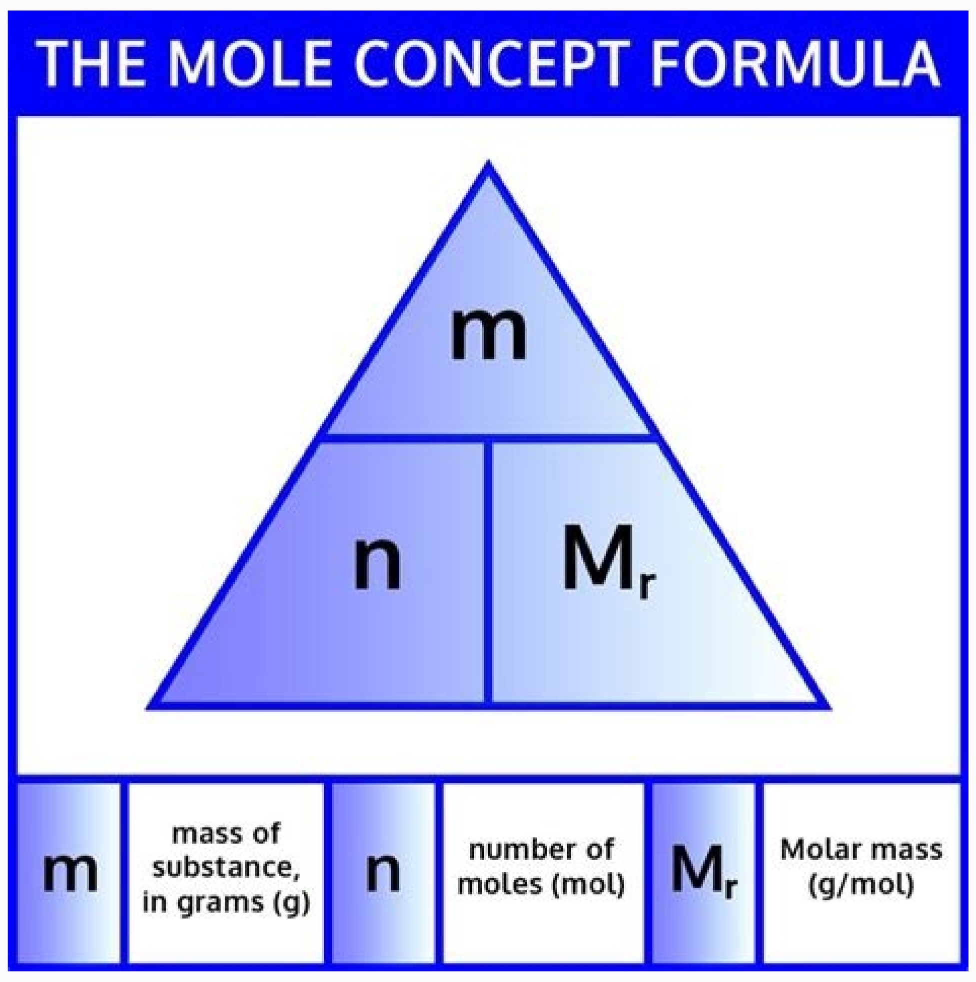

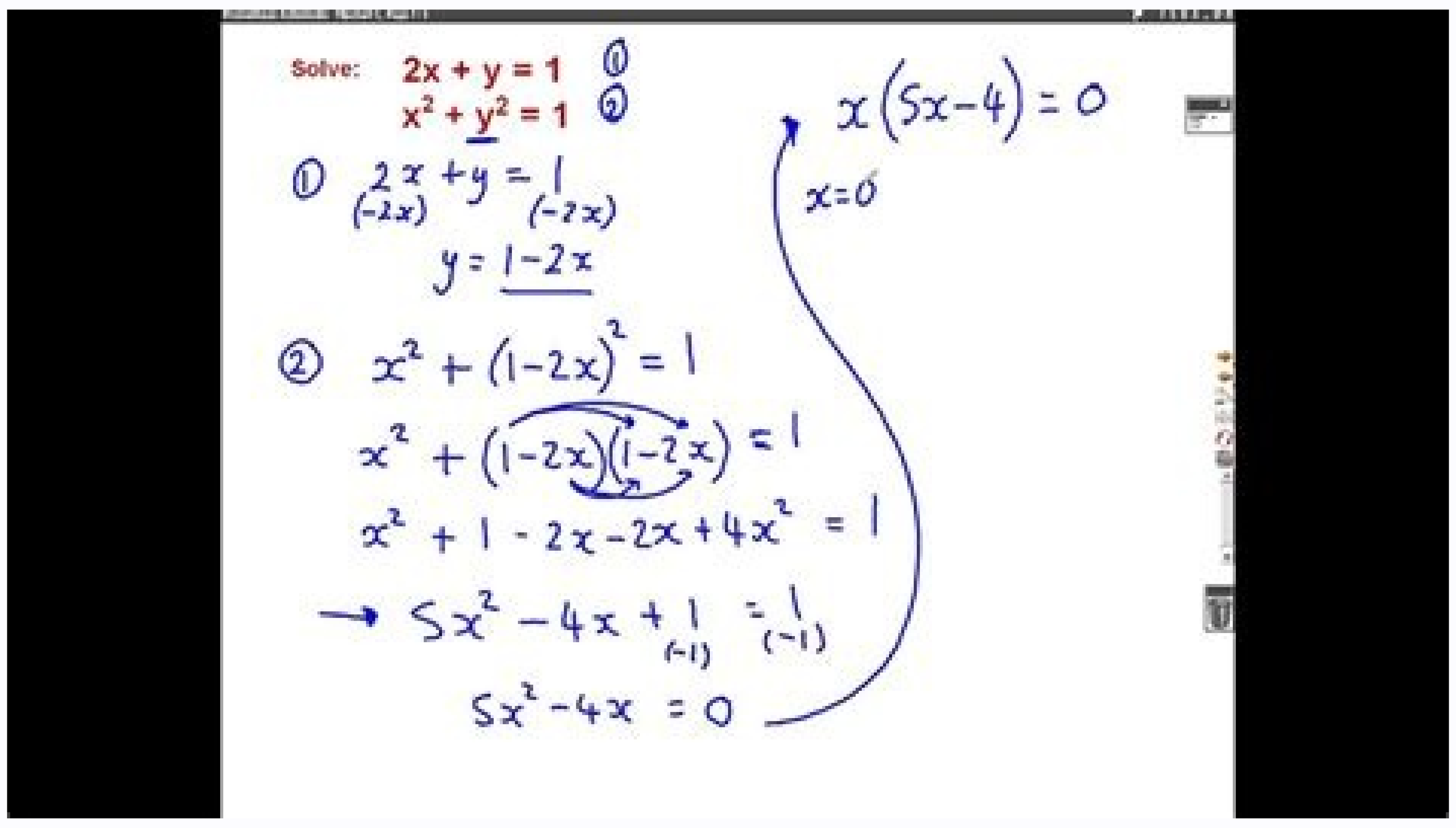

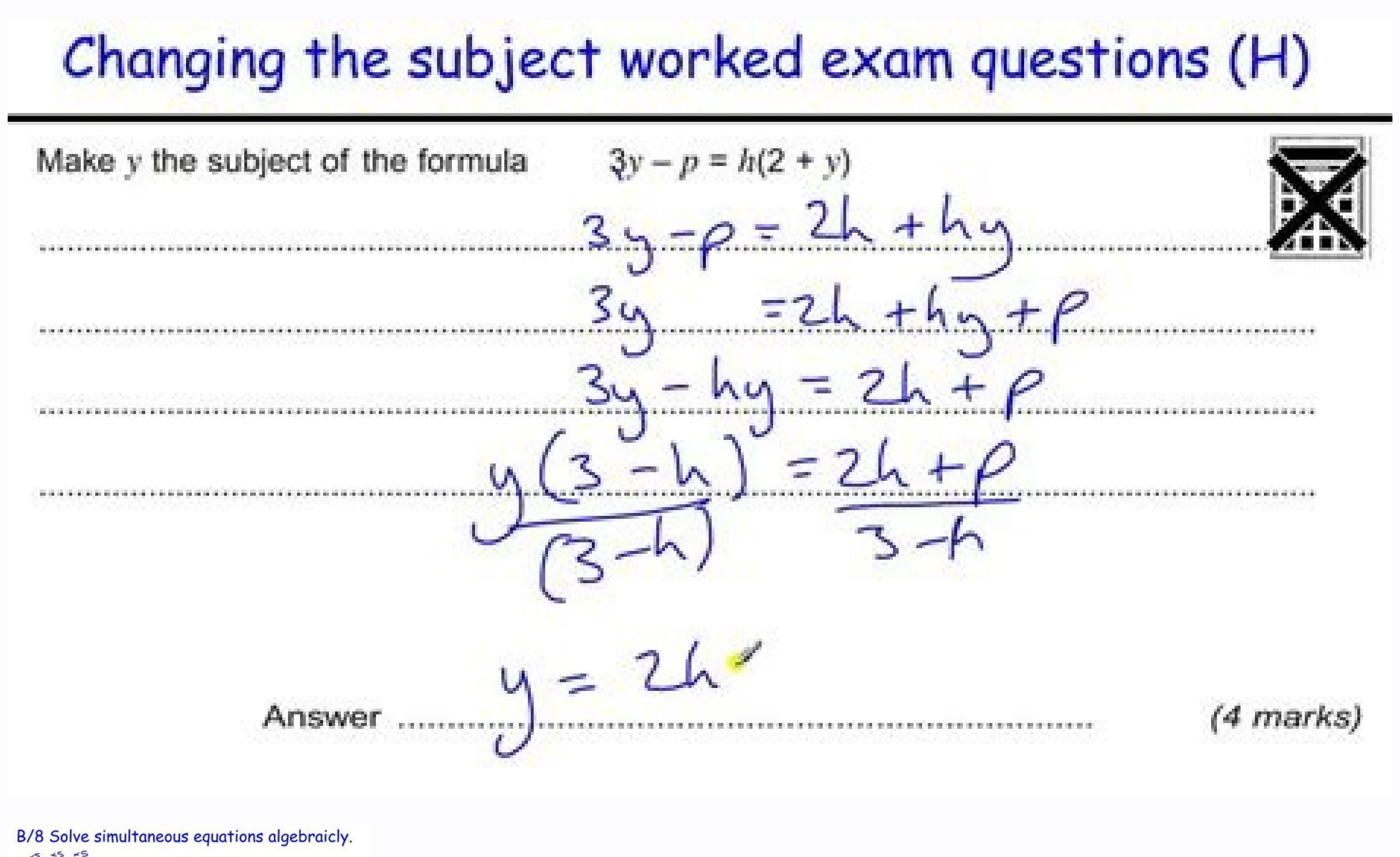

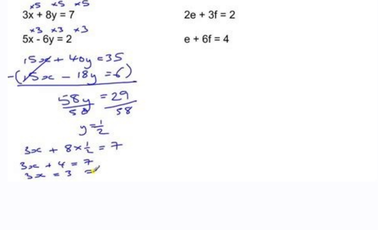

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Index of all my GCSE notes on acids, bases and salts All my GCSE Chemistry Revision notes This is a BIG website, you need to take time to explore it [SEARCH BOX] Use your mobile phone or ipad etc. in 'landscape' mode email doc brown Part 4. Some important REACTIONS of important ACIDS Acids react with a wide range of metals, oxides, hydroxides and carbonates to form salts in neutralisation reactions. The reactions of acids with metals, oxides, hydroxides, carbonates and hydrogencarbonates are described and lots of examples of word and symbol equations. Part 4 Describes and explains the reactions of common acids like hydrochloric acid, sulfuric acid and nitric acid with moderately reactive metals, metal oxides, metal hydroxides, metal carbonates and aqueous ammonia solution. What is formed in these reactions? Are the products of these reactions of any use? These revision notes on chemical reactions of acids e.g. sulfuric, hydrochloric and nitric acids, should prove useful for the new AQA chemistry, Edexcel chemistry & OCR chemistry GCSE (9–1, 9-5 & 5-1) science courses. EQUATION NOTE: The equations are often written three times: (i) word equation, (ii) balanced symbol equation without state symbols, and, (iii) with the state symbols (g), (l), (s) or (aq) to give the complete balanced symbol equation. 4. Some important reactions of Acids Acids are neutralised by reaction with metals, oxides, hydroxides or carbonates to form salts and other products. Apart from metals (which is an electron loss/gain redox reaction), the other reactants listed above are considered as bases, meaning they react by accepting a proton from an acid in forming the salt. Water soluble bases are known as alkalis. The reaction between acids and bases/alkalis like oxides, hydroxides and carbonates are classified as neutralisation reactions. The first part of the salt name is derived from the positive ion in the base e.g. sodium, magnesium, ammonium etc. The second part of the salt name is derived from the acid e.g. chloride (from hydrochloric acid), sulfate (from sulfuric acid), nitrate (from nitric acid) etc. You need to be able to predict the products of these reactions and know how to work out the formula of a salt given common ions. The ionic theory of neutralisation is described and explained in section 2. REACTION OF ACIDS WITH METALS General word equation: metal + acid ==> a salt + hydrogen The salt, and its name, depends on the metal and acid used in the reaction and the acid is neutralised in the process. e.g. the grey-silvery solid zinc dissolves in hydrochloric acid with effervescence to evolve hydrogen gas and leave a colourless solution of the salt zinc chloride. zinc + hydrochloric acid ==> zinc chloride + hydrogen Zn + 2HCl ==> ZnCl2 + H2 Zn(s) + 2HCl(aq) ==> ZnCl2(aq) + H2(g) (equation with state symbols) Its the same equation for many other Group 2 and Transition metals e.g. Mg, Ca and Fe, Co, Ni instead of zinc/Zn in the word/symbol equations Note that hydrochloric acid gives a chloride salt. Similarly magnesium + hydrochloric acid ==> magnesium chloride + hydrogen Mg + 2HCl ==> MgCl2 + H2 in pictures! iron + hydrochloric acid ==> iron(II) chloride + hydrogen aluminium + hydrochloric acid ==> aluminium chloride + hydrogen 2Al + 6HCl ==> 2AlCl3 + 3H2 (testing equation to balance) (you can add the state symbols as for the Zn + HCl reaction above) Illustrated above are two ways in which the zinc - hydrochloric reaction can be used to prepare a sample of hydrogen gas. A strip of magnesium ribbon dissolves with effervescence to evolve hydrogen gas and leave a colourless solution of the salt magnesium sulfate. magnesium + sulfuric acid ==> magnesium sulfate + hydrogen Mg + H2SO4 ==> MgSO4 + H2 Mg(s) + H2SO4(aq) ==> MgSO4(aq) + H2(g) (with state symbols) Note that sulfuric/sulfuric gives a sulfate/sulfate salt Similarly ... zinc + sulfuric acid ==> zinc sulfate + hydrogen Zn + H2SO4 ==> ZnSO4 + H2 Zn(s) + H2SO4(aq) ==> ZnSO4(aq) + H2(g) Instead of Mg or Zn, you can have Fe or Al e.g. iron + sulfuric acid ==> iron(II) sulfate + hydrogen Fe + H2SO4 ==> FeSO4 + H2 aluminium + sulfuric acid ==> aluminium sulfate + hydrogen 2Al + 3H2SO4 ==> Al2(SO4)3 + 3H2 (testing equation to balance!) However some metals give little or no reaction e.g. copper, therefore to make copper salts you need to react the acid with copper oxide or copper carbonate (see below in the following sections for the details of these reactions). You can test the gas given off with lit splint and the sound of a squeaky pop confirms it to be hydrogen gas AND therefore the original solid was a metal. 2H2 + O2 ==> 2H2O + heat energy! 2H2(g) + O2(g) ==> 2H2O(l) You might see the condensed water on side of test tube when doing this test. Note that nitric acid (HNO3) doesn't usually form hydrogen with a metal, instead you get nasty brown fumes of nitrogen dioxide! but you still get the metal nitrate salt .... so in general .... metal + nitric acid ==> metal nitrate salt + water + nitrogen oxides (maybe some hydrogen too) The reaction of metals with acids is a REDOX reaction and NOT an acid-base reaction. See also the REACTIVITY SERIES OF METALS page for the relative rate of reaction of metals with hydrochloric acid and sulfuric acid. Naming salts reminder - hydrochloric acid makes chloride salts, sulfuric/sulfuric acid makes sulfate/sulfate salts and nitric acid makes nitrate salts. This is the second part of the name, and the first part of the name (in most cases) is simply the metal name from the metal compound that reacted with the acid. However, sometimes you may need to add a Roman numeral in brackets to indicate the valency of that metal in a particular compound where a metal has a variable valency e.g. copper(II) .., iron(II) ... or an iron(III) salt ... etc. Advanced REDOX theory of the metal - acid reaction (this theory does NOT apply to any other reaction on this page because the reactions of acids with oxides, hydroxides or carbonates does NOT involve oxidation and reduction) Introduction to oxidation and reduction theory and application to REDOX reactions The reaction between a metal and an acid is technically what is called a REDOX reaction. This means the reaction takes place in two parts, an oxidation involving electron loss and a reduction involving electron gain. The metal atoms lose electrons to form positive ions (an oxidation - electron loss - the metal atoms are oxidised). The hydrogen ions gain electrons forming hydrogen gas molecules (reduction - electron gain - hydrogen ions are reduced). The two simultaneous changes occur on the surface of the metal where the aqueous positive hydrogen ions hit the surface and pinch electrons from the metal atoms and then metal ions pass into solution. These changes can be written as half equations and then combined to give the full redox ionic equation. e.g. for the reaction of magnesium with sulfuric acid, you can write (i) Mg(s) + H2SO4(aq) ==> MgSO4(aq) + H2(g) (ii) Mg(s) ===> Mg2+(aq) + 2e- (oxidation half equation, 2 electron loss from Mg atom) (iii) 2H+(aq) + 2e- ===> H2(g) (reduction half equation, 1 electron gain by 2 H+ ions) adding (ii) + (iii) gives the full (iv) the redox ionic equation ... (iv) Mg(s) + 2H+(aq) ===> Mg2+(aq) + H2(g) ... because the two electrons cancel each other out you don't see them in either the 'normal' equation or the ionic equation. The sulfate ion from the acid is the spectator ion, takes no part in the reaction and is not written into the ionic equation. You can write exactly the same sort of equations, whatever the acid and can also substitute Mg with e.g. Zn or Fe. e.g. zinc + hydrochloric acid ==> zinc chloride + hydrogen Zn(s) + 2HCl(aq) ==> ZnCl2(aq) + H2(g) the chloride ion Cl– is the spectator ion zinc + hydrogen ion ==> zinc ion + hydrogen The fully balanced symbol ionic equation is ... Zn(s) + 2H+(aq) ==> Zn2+(aq) + H2(g) Zinc atoms are oxidised to zinc ions by electron loss, so zinc is the reducing agent (electron donor). hydrogen ions are the oxidising agent (gaining the electrons) and are reduced to form hydrogen molecules by electron gain. Again, you can think of it as two 'half–reactions' Zn ==> Zn2+ + 2e– (the oxidation half equation, electron loss, zinc atom is oxidised to the zinc ion) 2H+ + 2e– ==> H2 (the reduction half equation, electron gain, hydrogen ions reduced to hydrogen molecules) again, the electron loss and gain cancel out, so you don't see the electrons in the ionic equation. Here, the chloride ion is the spectator ion and is not included in the ionic equation. The reaction of aluminium with hydrochloric acid or sulfuric acid would be normally given as: aluminium + hydrochloric acid ===> aluminium chloride + hydrogen 2Al + 6HCl ===> 2AlCl3 + 3H2 aluminium + sulfuric acid ===> aluminium sulfate + hydrogen 2Al + 3H2SO4 ===> Al2(SO4)3 + 3H2 I've omitted the state symbols, but still tricky equations to balance! The half reactions are the same because the chloride ions or sulfate ions are spectator ions. The oxidation half equation is: Al ===> Al3+ + 3e- (electron loss, aluminium atoms oxidised to aluminium ion) The reduction half equation is: 2H+ + 2e- ===> H2 (electron gain, hydrogen ions reduced to hydrogen molecules) To write the overall redox ionic equation by combining the two half equations is a little bit tricky! 2Al + 6H+ ===> 2Al3+ + 3H2 Note that the chloride ion (Cl-) or sulfate ions (SO42-) don't figure here, they are spectator ions, and don't take part in the reaction, its just the hydrogen ions that are involved from the acid - their formation is shown below. You can think of hydrochloric acid behaving as: HCl(aq) ===> H+(aq) + Cl-(aq) and for sulfuric acid: H2SO4(aq) ===> 2H+(aq) + SO42-(aq) because that is exactly what happens when you dissolve hydrogen chloride and sulfuric acid in water - they ionise and its the hydrogen ion that reacts with the metal surface. The chloride and sulfate ions simply become part of the salt solution formed in the reaction, but didn't actually chemically change like the metal atoms and hydrogen ions do. See also REDOX reactions and oxidation and reduction for more examples REACTION OF ACIDS WITH BASES - basic oxides and hydroxides These may be alkalis (soluble bases) or water insoluble bases Metal oxides and metal hydroxides are typical bases, they may be insoluble in water, or soluble in water to give an alkaline solution (a soluble base is called an alkali). An oxide formula just as an 'M' + 'O', a hydroxide formula has an 'M' + 'OH' in various simple proportions. The general word equation for this classic 'neutralisation' reaction is alkali (soluble/insoluble base) + acid ==> salt + water i.e. the acid is neutralised in the process. Reactions of acids with soluble bases (alkalis, usually soluble metal hydroxides) metal hydroxide + acid ==> a salt + water The acid is neutralised in the process e.g. sodium hydroxide + hydrochloric acid ==> sodium chloride + water NaOH + HCl ==> NaCl + H2O NaOH(aq) + HCl(aq) ==> NaCl(aq) + H2O(l) (equation with state symbols) in pictures! Other examples of neutralising alkalis (soluble bases) potassium hydroxide + hydrochloric acid ==> potassium chloride + water potassium hydroxide + hydrobromic acid ==> potassium bromide + water KOH(aq) + HBr(aq) ==> KBr(aq) + H2O(l) (equation with state symbols) Its the same for any Group 1 Alkali Metal hydroxide e.g. LiOH, NaOH, RbOH etc. and any other Group 7 halogen acids e.g. HCl hydrochloric acid. calcium hydroxide + hydrochloric acid ==> calcium chloride + water Ca(OH)2 + 2HCl ==> CaCl2 + 2H2O All solutions involved are colourless but the reaction can be monitored using universal indicator or a pH meter. Soluble bases i.e. here, soluble oxides or hydroxides, are called alkalis. This neutralisation reaction is used in ... or sodium hydroxide + sulfuric acid ==> sodium sulfate + water 2NaOH + H2SO4 ==> Na2SO4 + 2H2O 2NaOH(aq) + H2SO4(aq) ==> Na2SO4(aq) + 2H2O(l) (equation with state symbols) Its the same equation for any Group 1 Alkali Metal hydroxide e.g. LiOH, KOH etc. e.g. sodium hydroxide + nitric acid ==> sodium nitrate + water NaOH + HNO3 ==> NaNO3 + 2H2O NaOH(aq) + HNO3(aq) ==> NaNO3(aq) + H2O(l) potassium hydroxide + nitric acid ==> potassium nitrate + water KOH + HNO3 ==> KNO3 + H2O KOH(aq) + HNO3(aq) ==> KNO3(aq) + H2O(l) Another soluble base is ammonia, and here are the equations for its neutralisation with the three most common mineral acids you come across in the school/college laboratory. ammonia + nitric acid ==> ammonium nitrate NH3 + HNO3 ==> NH4NO3 NH3(aq) + HNO3(aq) ==> NH4NO3(aq) ammonia + hydrochloric acid ==> ammonium chloride NH3 + HCl ==> NH4Cl NH3(aq) + HCl(aq) ==> NH4Cl(aq) ammonia + sulfuric acid ==> ammonium sulfate 2NH3 + H2SO4 ==> (NH4)2SO4 2NH3(aq) + H2SO4(aq) ==> (NH4)2SO4(aq) You should have noticed that with hydrochloric acid (HCl) and nitric acid (HNO3), there is one hydrogen (H) that is replaced by the metal ion, and in the case of sulfuric acid (H2SO4), there are two protons replaced by a metal ion or ions. This is because each HCl or HNO3 molecule produces one hydrogen ion (H+) in water, and each sulfuric acid molecule produces two hydrogen ions in water. There is one common mineral acid that potentially can produces three protons per acid molecule. This is phosphoric(V) acid, H3PO4, commonly known as just phosphoric acid. Therefore, phosphoric acid, will react with three times as much sodium hydroxide when the solutions are mixed to form the salt trisodium phosphate(V) (or simply 'sodium phosphate'). phosphoric acid + sodium hydroxide ==> sodium phosphate + water H3PO4 + 3NaOH ==> Na3PO4 + 3H2O H3PO4(aq) + 3NaOH(aq) ==> Na3PO4(aq) + 3H2O(l) Reactions of acids with insoluble bases (metal oxides and metal hydroxides) insoluble base + acid ==> salt + water (note: oxides that react with acids to form salts are known as 'basic oxides') The acid is neutralised in the process e.g. e.g. metal oxide + acid ==> salt* + water Black copper (ii) oxide dissolves in colourless sulfuric acid to give a blue solution of copper(II) sulfate from which the salt can be crystallised. copper(II) oxide + sulfuric acid ==> copper(II) sulfate + water CuO + H2SO4 ==> CuSO4 + H2O CuO(s) + H2SO4(aq) ==> CuSO4(aq) + H2O(l) Note that insoluble bases, like these insoluble oxides, although reacting with acids to form salts, are NOT alkalis. Instead of copper/Cu, you can have magnesium/Mg, zinc/Zn or nickel/Ni in the word/symbol equations. This neutralisation reaction is used in salt preparations method b for details. Apart from copper compounds, all solutions involved here are colourless and all the salts form colourless crystal if the solution is carefully evaporated to cause crystallisation. calcium hydroxide + hydrochloric acid ==> calcium chloride + water Ca(OH)2 + 2HCl ==> CaCl2 + 2H2O Ca(OH)2(s) + 2HCl(aq) ==> CaCl2(aq) + 2H2O(l) calcium hydroxide + sulfuric acid ==> calcium sulfate + water Ca(OH)2 + H2SO4 ==> CaSO4 + 2H2O Ca(OH)2(s) + H2SO4(aq) ==> CaSO4(aq/s) + 2H2O(l) This reaction soon slows down because calcium sulfate isn't very soluble in water. Other examples ... sulfuric acid + zinc oxide ==> zinc sulfate + water H2SO4 + ZnO ==> ZnSO4 + H2O magnesium oxide + nitric acid ==> magnesium nitrate + water MgO +

2HNO3 ==> Mg(NO3)2 + H2O calcium oxide + nitric acid ==> calcium nitrate + water CaO + 2HNO3 ==> Ca(NO3)2 + H2O magnesium hydroxide + hydrochloric acid ==> magnesium chloride + water Mg(OH)2(s) + 2HCl(aq) ==> MgCl2(aq) + 2H2O(l) and ... magnesium hydroxide + nitric acid ===> magnesium nitrate + water Mg(OH)2(s) + 2HNO3(aq) ==> Mg(NO3)2(aq) + 2H2O(l) in pictures Similar for many other Group 2 and Transition metal hydroxides e.g. Ca, Sr, Ba and Fe, Co, Ni, Cu, Zn instead of Mg zinc oxide + hydrochloric acid ==> zinc chloride + water ZnO(s) + 2HCl(aq) ==> ZnCl2(aq) + H2O(l) Similar for many other Group 2 and Transition metal oxides e.g. Mg, Ca, Ba and Co, Ni, Cu instead of Zn copper(II) oxide + hydrochloric acid ==> copper(II) chloride + water CuO(s) + 2HCl(aq) ==> CuCl2(aq) + H2O(l) zinc oxide + sulfuric acid ==> zinc sulfate + water ZnO + H2SO4 ==> ZnSO4 + H2O ZnO(s) + H2SO4(aq) ==> ZnSO4(aq) + H2O(l) and finally, a bit more tricky! aluminium oxide + sulfuric acid ===> aluminium sulfate + water Al2O3 + 3H2SO4 ===> Al2(SO4)3 + 3H2O Apart from copper compounds, all solutions involved here are colourless and all the salts form colourless crystal if the solution is carefully evaporated to cause crystallisation. Naming salts reminder - hydrochloric acid makes chloride salts, sulfuric/sulfuric acid makes sulfate/sulfate salts and nitric acid makes nitrate salts. Reactions of acids with soluble/insoluble carbonates and hydrogencarbonates (also bases, most insoluble) Contrary to what some textbooks may say, but often do not point out, all carbonates are bases and react with acids to form salts, the difference in reaction with alkalis or insoluble bases is that carbon dioxide gas is evolved. On adding a solid carbonate or hydrogencarbonate to an acid you see effervescence from carbon dioxide gas and the general word equation is ... carbonate/hydrogencarbonate + acid ==> a salt + water + carbon dioxide The obvious extra observation compared to the reaction with oxides or hydroxides is the production of a gas, but the acid is still neutralised in the process to form a salt e.g. The white solid calcium carbonate (limestone) dissolves in dilute hydrochloric acid to form a colourless solution of calcium chloride and colourless carbon dioxide gas. This is a typical reaction between an insoluble carbonate and a dilute acid solution. calcium carbonate + hydrochloric acid ==> calcium chloride + water + carbon dioxide CaCO3 + 2HCl ==> CaCl2 + H2O + CO2 CaCO3(s) + 2HCl(aq) ==> CaCl2(aq) + H2O(l) + CO2(g) (equation with state symbols) Illustrated above are two ways in which the limestone chips (calcium carbonate) - hydrochloric acid reaction can be used to prepare a sample of carbon dioxide gas. This neutralisation reaction is used in salt preparations (see method b for details). Apart from copper compounds, all solutions involved here are colourless and all the salts form colourless crystal if the solution is carefully evaporated to cause crystallisation. The dark turquoise-green solid copper(II) carbonate dissolves in hydrochloric acid to form a greeny-blue solution of copper(II) chloride and effervescence from the carbon dioxide formed. Another reaction between an insoluble carbonate and a dilute acid solution. copper(II) carbonate + hydrochloric acid ==> Copper(II) chloride + water + carbon dioxide CuCO3 + 2HCl ==> CuCl2 + H2O + CO2 CuCO3(s) + 2HCl(aq) ==> CuCl2(aq) + H2O(l) + CO2(g) and with sulfuric acid a blue solution of copper(II) sulfate is formed and with nitric acid a blue solution of copper(II) nitrate - word and full symbol equations below. copper(II) carbonate + sulfuric acid ==> Copper(II) sulfate + water + carbon dioxide CuCO3 + H2SO4 ==> CuSO4 + H2O + CO2 CuCO3(s) + H2SO4(aq) ==> CuSO4(aq) + H2O(l) + CO2(g) in pictures! copper(II) carbonate + nitric acid ==> Copper(II) nitrate + water + carbon dioxide CuCO3 + 2HNO3 ==> Cu(NO3)2 + H2O + CO2 CuCO3(s) + 2HNO3(aq) ==> CuSO4(aq) + H2O(l) + CO2(g) Similar equations for other carbonates calcium carbonate CaCO3, to make three salts - calcium chloride/nitrate (calcium sulfate is not very soluble) iron(II) carbonate FeCO3, to make three salts - iron(II) chloride/sulfate/nitrate magnesium carbonate MgCO3, to make three salts - magnesium chloride/sulfate/nitrate manganese(II) carbonate MnCO3, to make three salts - manganese(II) chloride/sulfate/nitrate zinc carbonate ZnCO3, to make three salts - zinc chloride/sulfate/nitrate lead(II) carbonate PbCO3, only nitric acid to make lead(II) nitrate Similarly, but forming colourless solutions from other white insoluble solid carbonates ... magnesium carbonate + hydrochloric acid ==> magnesium chloride + water + carbon dioxide MgCO3 + 2HCl ==> MgCl2 + H2O + CO2 MgCO3(s) + 2HCl(aq) ==> MgCl2(aq) + H2O(l) + CO2(g) or calcium carbonate + nitric acid ==> calcium nitrate + water + carbon dioxide CaCO3 + 2HNO3 ==> Ca(NO3)2 + H2O + CO2 CaCO3(s) + 2HNO3(aq) ==> Ca(NO3)2(aq) + H2O(l) + CO2(g) Its the same equation for many other Group 2 and Transition metals e.g. Mg, Sr and Co, Ni, Cu Test for carbon dioxide gas - it gives a white precipitate of calcium carbonate (cloudiness) when bubbled into limewater (calcium hydroxide solution). Ca(OH)2 + CO2 ==> CaCO3 + H2O Ca(OH)2(aq) + CO2(g) ==> CaCO3(s) + H2O(l) Note: Using sulfuric acid and calcium carbonate you don't get much of a fizz! because the calcium sulfate salt formed, is not very soluble, and coats the remaining calcium carbonate inhibiting the reaction! This will happen with any reaction between an acid and a water insoluble reactant which forms an insoluble solid product! magnesium carbonate + sulfuric acid ==> magnesium sulfate + water + carbon dioxide MgCO3 + H2SO4 ==> MgSO4 + H2O + CO2 MgCO3(s) + H2SO4(aq) ==> MgSO4(aq) + H2O(l) + CO2(g) AND six equations for sodium carbonate and sodium hydrogencarbonate. Both are actually soluble in water, but in each case you are likely to add the white solid directly into the acid where it dissolves to give a colourless solution of the colourless salt with the evolution of carbon dioxide gas (fizzing - effervescence) e.g.... sodium hydrogencarbonate + hydrochloric acid ==> sodium chloride + water + carbon dioxide NaHCO3 + HCl ==> NaCl + H2O + CO2 NaHCO3(s) + HCl(aq) ==> NaCl(aq) + H2O(l) + CO2(g) Note that hydrogencarbonate is commonly known as 'bicarbonate'. Its the same for any Group 1 Alkali Metal hydrogencarbonate e.g. for Li, K etc. instead of Na sodium hydrogencarbonate + nitric acid ==> sodium nitrate + water + carbon dioxide NaHCO3 + HNO3 ==> NaNO3 + H2O + CO2 NaHCO3(s) + HNO3(aq) ==> NaNO3(aq) + H2O(l) + CO2(g) sodium carbonate + hydrochloric acid ==> sodium chloride + water + carbon dioxide Na2CO3 + 2HCl ==> 2NaCl + H2O + CO2 Na2CO3(s) + 2HCl(aq) ==> 2NaCl(aq) + H2O(l) + CO2(g) sodium hydrogencarbonate + sulfuric acid ==> sodium sulfate + water + carbon dioxide 2NaHCO3 + H2SO4 ==> Na2SO4 + 2H2O + 2CO2 2NaHCO3(s) + H2SO4(aq) ==> Na2SO4(aq) + 2H2O(l) + 2CO2(g) sodium carbonate + sulfuric acid ==> sodium sulfate + water + carbon dioxide Na2CO3 + H2SO4 ==> Na2SO4 + H2O + CO2 Na2CO3(s) + H2SO4(aq) ==> Na2SO4(aq) + H2O(l) + CO2(g) sodium carbonate + nitric acid ==> sodium nitrate + water + carbon dioxide Na2CO3 + 2HNO3 ==> 2NaNO3 + H2O + CO2 Na2CO3(s) + 2HNO3(aq) ==> 2NaNO3(aq) + H2O(l) + CO2(g) phosphoric acid + sodium carbonate ==> sodium phosphate + water + carbon dioxide 2H3PO4 + 3Na2CO3 ==> 2Na3PO4 + 3H2O + 3CO2 2H3PO4(aq) + 3Na2CO3(aq) ==> 2Na3PO4(aq) + 3H2O(l) + 3CO2(g) That's a really tricky equation to balance!!! You can test the gas given off with limewater and the formation of a white precipitate confirms it to be carbon dioxide AND therefore the original solid was a carbonate or hydrogencarbonate. For more reactions of acids and carbonates see Limestone Chemistry page Naming salts reminder - hydrochloric acid makes chloride salts, sulfuric/sulfuric acid makes sulfate/sulfate salts and nitric acid makes nitrate salts. REACTIONS OF ACIDS WITH AMMONIA Ammonia gas is very soluble in water to form an alkaline solution that can be neutralised by acids to form ammonium salts. All solutions involved here are colourless and all the salts form colourless crystal if the solution is carefully evaporated to cause crystallisation. ammonia + acid ==> ammonium salt Note that no water is formed and see also note (c) below. e.g. (i) ammonia + hydrochloric acid ==> ammonium chloride NH3 + HCl ==> NH4Cl NH3(aq) + HCl(aq) ==> NH4Cl(aq) or (ii) ammonia + nitric acid ==> ammonium nitrate NH3 + HNO3 ==> NH4NO3 NH3(aq) + HNO3(aq) ==> NH4NO3(aq) or (iii) ammonia + sulfuric acid ==> ammonium sulfate 2NH3 + H2SO4 ==> (NH4)2SO4 2NH3(aq) + H2SO4(aq) ==> (NH4)2SO4(aq) NOTE that (a) NH3 is used in equations, NOT NH4OH which doesn't exist! Therefore, theoretically, no water is produced when ammonia solution is neutralised with acids, BUT it is still a neutralisation reaction because ammonia is a base/alkali. (b) All these ammonium salts are colourless crystalline solids - formed if the water is carefully evaporated salt preparation method (a) for details and on ammonium salts page for details. (c) Reactions (ii) and (iii) are used to make fertiliser salts - see ammonia chemistry and uses. Naming salts reminder - hydrochloric acid makes chloride salts, sulfuric acid makes sulfate salts and nitric acid makes nitrate salts. NOTE (a): As already mentioned, and to summarise, the name of the particular salt formed depends on (i) the metal name, which becomes the first part of salt name, and (ii) the acid e.g. H2SO4 sulfuric acid on neutralisation makes a ... sulfate; HCl hydrochloric acid makes a ... chloride; HNO3 nitric acid makes a ... nitrate etc. NOTE (b): There is a list of compound formulae and their solubility in section 8. The first part of the salt name is ammonium derived from ammonia (with metals or their compounds the metal retains its original name), but the second part of the salt name is always derived from the acid as in NOTE (a) above. NOTE (c): Ammonia is an alkaline gas that is very soluble in water. It is a weak alkali or soluble base and is readily neutralised by acids in solution to form ammonium salts which can be crystallised on evaporating the resulting solution. Sometimes the equations are written with the 'fictitious' 'ammonium hydroxide' e.g. NH4OH(aq) + HCl(aq) ==> NH4Cl(aq) + H2O(l) but this is not considered to be a correct representation of this neutralisation reaction these days! NOTE (d): An extensive structured question on acid reaction equations. Appendix How to work out the formula of a salt given the ions In the formula of the salt the total positive charge must equal the negative charge i.e. the salt must be overall electrically neutral. A list of common positive ions (cations e.g. from the base) and negative ions (e.g. from the acid) is given on the right. In the examples below of salt formulae, the derived formula are shown in 'molecular' formula style, but they are actually ionic compounds, so the ionic formula is also shown. For potassium chloride: 1 of K+ balances 1 of Cl- because 1 x 1 = 1 x 1 gives KCl or K+Cl- For magnesium chloride: 1 of Mg2+ balances 2 of Cl- because 1 x 2 = 2 x 1 gives MgCl2 or Mg2+(Cl-)2 For iron(III) chloride: 1 of Fe3+ balances 3 of Cl- because 1 x 3 = 3 x 1 gives FeCl3 or Fe3+(Cl-)3 For sodium sulfate: 2 of Na+ balances 1 of SO42- because 2 x 1 = 1 x 2 gives Na2SO4 or (Na+)2SO42- For calcium nitrate: 1 of Ca2+ balances 2 of NO3- because 1 x 2 = 2 x 1 gives Ca(NO3)2 or Ca2+(NO3-)2 For iron(III) sulfate: 2 of Fe3+ balances 3 of SO42- because 2 x 3 = 3 x 2 gives Fe2(SO4)3 or (Fe3+)2(SO42-)3 For more on equations and formulae see: How to write equations, work out formula and name compounds GCSE/IGCSE Acids & Alkalis revision notes sub–index: Index of all pH, Acids, Alkalis, Salts Notes 1. Examples of everyday acids, alkalis, salts, pH of solution, hazard warning signs : 2. pH scale, indicators, ionic theory of acids–alkali neutralisation : 4. Reactions of acids with metals/oxides/hydroxides/carbonates, neutralisation reactions : 5. Reactions of bases–alkalis like ammonia & sodium hydroxide : 6. Four methods of making salts : 7. Changes in pH in a neutralisation, choice and use of indicators : 8. Important formulae of compounds, salt solubility and water of crystallisation : 10. 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