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Chapter 1 : Iron and sulfur reaction- Learn Chemistry

Notes on Metallurgical Analysis, Arranged for Students in Metallurgical Chemistry Selected Methods for the Analysis of Iron and Steel and of the Materials Used in Their Manufactures, Including the Analysis of Gases, Fuels, Water for Boiler Supply, Etc by Nathaniel Wright Lord.

The electrode potential chart above highlights the values for various oxidation states of iron. Hydrogen gas is evolved and it is a redox reaction. Note that the lower oxidation state of iron is formed, since neither acid is a strong oxidising agent. The pale green salts FeCl2. However, they are readily oxidised by dissolved oxygen to form iron III compounds more on this later. White anhydrous iron II chloride can be made by passing hydrogen chloride gas over heated iron. Iron III chloride is a brown covalently bonded, relatively volatile chloride. It exists in the solid form as a covalent molecular lattice of the dimer Fe2Cl6. So, strictly speaking the equation should be written as: The exothermic nature of the reaction may or may not be seen? The dimer molecules are present in the brown solid. When iron wool is heated with iodine there is little reaction, a small amount of iron II iodide is formed. This is theoretically predictable from the half-cell electrode potentials. The iron III chloride reacts very exothermically with water to give pungent acrid fumes of hydrogen chloride anhydrous aluminium chloride is made in the same way and behaves with water in the same way! Hence the need for dry conditions in their preparation is illustrated below. Its also a very good idea to vent the excess chlorine away safely! It has an octahedral shape and a co-ordination number of 6 from 6 unidentate ligands. What you normally see is the yellowâ€"light brownâ€"orange coloured complex ion formed from proton transfer to water giving a hydroxoâ€"complex ion see equation below. This proton transfer process can continue in higher pH media to give the iron III hydroxide precipitate see later and accounts for why iron III salt solutions are acidic. There is no further reaction with excess of either i. All are acidâ€"base reactions and not redox reactions except that iron II compounds can be readily oxidised to iron III compounds by the oxygen in air.. NH3 is a weak base but slightly ionises in water to give sufficient hydroxide ions to give the precipitates. Aqueous sodium carbonate is weakly alkaline and gives the hydroxide ppts. Again, theses are all acidâ€"base reactions and not redox changes. The iron II ion probably gives a mixture of the hydroxide see above and carbonate too a basic carbonate? The acidity of the hydrated iron III ion makes it react with the carbonate ion. The chlorine water itself is a very pale green, and changes to the colourless chloride ion, so the colour change associated with the oxidation state change of iron II to iron III is quite clearly seen. Note that chlorine is a powerful enough oxidising agent to oxidise iron II ion to the iron III ion, BUT iodine is not a strong enough oxidising agent to achieve this. It is in fact the iron III ion that will oxidise the iodide ion, rather than the reverse. So, crossâ€"check the reaction the oxidation of iodide ions by iron III ions described below.

Chapter 2 : Notes on the chemistry of iron

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Chemistry in its element: End promo Chris Smith Hello, this week we turn to one of the most important elements in the human body. There are iron man challenges, iron fisted leaders and those said to have iron in the soul. This change from the relatively plentiful and soluble FeII, took a heavy toil on almost everything alive at the time. Surviving terrestrial and ocean-dwelling microbes developed soluble siderophore molecules to regain access to this plentiful, but otherwise inaccessible essential resource, which used hydroxamate or catechol chelating groups to bring the FeIII back into solution. Eventually higher organisms including animals, evolved. And animals used the energy of oxygen recombining with the hydrocarbons and carbohydrates in plant life to enable motion. Iron was essential to this process. But no animal, however, has been able to adequately deal, in the long run - meaning eighty year life spans - with the fact that iron is essential for the conversion of solar energy to movement, but is virtually insoluble in water at neutral pH, and, even worse, is toxic. Systems have evolved to maintain iron in specific useful and safe configurations - enzymes which utilize its catalytic powers, or transferrins and haemosiderins, which move it around and store it. But these are not perfect. Sometimes iron atoms are misplaced, and there are no known systems to recapture iron that has precipitated inside of a cell. Neurons sprout thousands of processes during their existence - reaching out to form networks of connections to other neurons. During development of the adult human brain a large percentage of cells are completely eliminated, and some new ones are added. It is a learning process. But once an area of the brain is up and running, there is nothing that can be done biologically, if a large number of its cells stop working for any reason. And the slow creep of precipitating iron over many decades is perhaps most often that reason. In less sophisticated tissues, like the liver, new stem cells can be activated, but in the brain, trained, structurally complex, interconnected neurons are needed, with thousands of projections that are accumulated over a lifetime of learning. This same basic mechanism can result in a variety of diseases. There are twenty or thirty proteins that that deal with iron in the brain - holding iron and passing it from place to place. Every new individual endowed with a new set of chromosomes is endowed with a new set of these proteins. Some combinations will be better than others and some will be dangerous individually and collectively. A mutation in a gene that codes for one of these proteins could disrupt its function - allowing iron atoms to become lost. These atoms that have been lost from the chemical groups that hold them will not always be safely returned to some structure like transferrin or haemoferritin. Some of them will react with water and be lost forever. They are piling up in the unlucky cell types that were the designated locations for expression of the most iron-leaky proteins. And oxides of iron are not just taking up critical space. Iron is very reactive. The infamous "Reactive Oxygen Species" which have been suspected of causing so many age related illnesses may just derive from various forms of iron. It is time for specialists trained in chemistry, and with an eye to the chemistry of iron, to pay some attention to neurodegenerative disease. Next time on Chemistry in its Element Johnny Ball will tell the story of Marie Curie and the element that she discovered and then named after her homeland. Johnny Ball Pitchblende, a uranium bearing ore, seemed to be far too radio active than could be accounted for by the uranium. They sieved and sorted by hand ounce by ounce through tons of pitchblende in a drafty, freezing shed, before eventually tiny amounts of polonium were discovered. End promo Help text not available for this section currently Video.

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Chapter 3 : Chemistry of Iron - Chemistry LibreTexts

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In either process the oxidising agent gets reduced. In either case the reducing agent gets oxidised. The more reactive metal magnesium always displaces the less reactive metal iron. Displacement reactions involving metals and metal ions are electron transfer reactions. In metal â€" soluble metal salt displacement reactions, the metal atom always loses electrons oxidation and the metal ion always gains electrons reduction. Exactly the same reaction occurs if you add iron filings to copper sulfate solution. A brown precipitate of copper forms on the surface of the iron filings and the blue colour fades as the less reactive copper is displaced by the more reactive iron. The iron sulfate formed in solution is a very pale green so the final solution is almost colourless if excess iron is added. Thirdly, adding magnesium to blue copper II sulfate solution, the blue colour fades as colourless magnesium sulfate is formed and brown bits of copper metal form a precipitate around the magnesium ribbon: 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 oxidation. The hydrogen ions gain electrons to form hydrogen gas molecules reduction. The two simultaneous changes occur on the surface of the metal where the positive hydrogen ions hit the metal surface and pinch electrons from the metal and so metal ions pass into solution e. The more reactive metal copper always displaces the less reactive metal sliver. The original chloride ions are effectively spectator ions. The full balanced symbol ionic equation is Potassium manganate VII is a powerful oxidising agent and an intense purple colour in water due to the MnO4â€" ion. Potassium iodide is a colourless salt dissolving in water to form a colourless solution. If it is oxidised e. The use of Roman Numerals in names: This indicates what is called the oxidation state of an atom in a molecule or ion. It is easy to follow for simple metal ions because it equals the charge on the ion e.

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Chapter 4 : Periodic Table of Elements: Iron - Fe (calendrierdelascience.com)

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Iron powder is preferred to iron filings. If fine sulfur powder is mixed with iron filings, it is difficult to obtain a consistent mix, because the two solids can separate. Roll sulfur or flowers of sulfur should be finely powdered using a pestle and mortar. Place a small plug of mineral wool in the mouth of each ignition tube. After the experiment, the iron II sulfide is low hazard and can be discarded into the refuse. Do this by weighing out 7 g of iron powder and 4 g of finely powdered sulfur onto separate pieces of filter paper or use weighing boats. Mix the two powders by pouring repeatedly from one piece of paper to the other until a homogeneous mixture by appearance is obtained. Demonstrate that iron can be separated from the mixture by physical means. Do this by wrapping the end of a small bar magnet in a paper tissue or cling film, and dipping it into a teaspoon-sized heap of the mixture on a watch glass. The iron will be attracted, but the sulfur remains on the watch glass. Clamp the test-tube as shown in the diagram. Heat until an orange glow is seen inside the test-tube. Let the students see that the glow continues and moves steadily through the mixture. At this point the students could carry out their own small-scale version of the reaction. The test-tube can be broken open using a pestle and mortar. It is advisable to wear protective gloves. However, this is not always successful. It has been suggested that using a very weak magnet is advisable. Bunsen burners should then be turned off. It may be sensible to get the students to place all their used reaction tubes onto one heat resistant mat set aside for this purpose e. Teaching notes On heating the reaction mixture, the sulfur melts and reacts with the iron exothermically to form iron II sulfide. The mineral wool plug in the mouth of the test-tube prevents sulfur vapour escaping and possibly catching fire. If, despite all precautions, the sulfur vapour does ignite, students must be trained to extinguish it by placing a damp rag firmly over the mouth of the tube. The signs that a chemical reaction occurs are: The two solids are mixed and heated in a test-tube or ignition tube. The reaction can be used to illustrate elements, mixtures and compounds. This collection of over practical activities demonstrates a wide range of chemical concepts and processes. Each activity contains comprehensive information for teachers and technicians, including full technical notes and step-by-step procedures.

Rusting of iron consists of the formation of hydrated oxide, Fe OH 3 or FeO OH , and is an electrochemical process which requires the presence of water, oxygen and an electrolyte - in the absence of any one of these rusting does not occur to any significant extent. The process is complex and will depend in detail on the prevailing conditions, for example, in the presence of a small amount of O2 the anodic oxidation will be: If both water and air are present, then the corrosion can be severe with oxygen now as the oxidant the anodic oxidations: The presence of an electrolyte is required to provide a pathway for the current and, in urban areas, this is commonly iron II sulfate formed as a result of attack by atmospheric SO2 but, in seaside areas, airborne particles of salt are important. The anodic oxidation of the iron is usually localized in surface pits and crevices which allow the formation of adherent rust over the remaining surface area. The illustration above shows 2 nails immersed in an agar gel containing phenolphthalein and [Fe CN 6] The nails can be seen to have started to corrode since the Prussian blue formation indicates the formation of Fe II the Anodic sites which correspond to the end of the nails and the bend in the middle. The phenolphthalein change to pink in presence of base shows the build up of OH- and shows that essentially the whole length of the nail is acting as the cathode. Eventually the lateral extension of the anodic area undermines the rust to produce loose flakes. Moreover, once an adherent film of rust has formed, simply painting over gives but poor protection. This is due to the presence of electrolytes such as iron II sulfate in the film so that painting merely seals in the ingredients for anodic oxidation. It then only requires the exposure of some other portion of the surface, where cathodic reduction can take place, for rusting beneath the paint to occur. The protection of iron and steel against rusting takes many forms, including: Another method uses sacrificial anodes, most usually Mg or Zn which, being higher than Fe in the electrochemical series, are attacked preferentially. In fact, the Zn coating on galvanized iron is actually a sacrificial anode. One level of rust prevention occurs through a purely mechanical method since it is more difficult for water and oxygen to reach the iron. Even if the layer becomes somewhat worn though another reason corrosion is inhibited is that the anodic processes are affected. Foodstuffs are often distributed in "tin cans" and it has generally been easier to coat the iron with a layer of tin than with zinc. Another benefit is that tin is less reactive then zinc so does not react as readily with the contents. Another technique is to treat the iron surface with dichromate solution.

Chapter 6 : Effect of Metal Coupling on the Rusting of Iron - Chemistry Project - Notes For Free

The chemistry of iron is dominated by the +2 and +3 oxidation states i.e. iron(II) and iron(III) complexes e.g. with the ligands cyanide. Periodic Table - Transition Metals - Iron Chemistry - Doc Brown's Chemistry Revising Advanced Level Inorganic Chemistry Periodic Table Revision Notes.

Contributors Iron, which takes its English name from the old Anglo-Saxon and its symbol from the Latin, ferrum, was identified and used in prehistoric times. Introduction In its pure form, iron is a silvery-white metal, distinguished by its ability to take and retain a magnetic field, and also dissolve small amounts of carbon when molten thus yielding steel. The oxides are reduced to pure iron. In addition to hardening iron by adding small amounts of carbon and also some other metals to the molten iron, iron castings or forgings can be heat-treated to take advantage of the various physical properties of the different solid phases of iron. Pure iron reacts readily with oxygen and moisture in the environment and corrodes destructively. Even alloys such as steel need protection by painting or some other coating to prevent structural failure over time. Iron as Catalyst The Haber Process combines nitrogen and hydrogen into ammonia. The nitrogen comes from the air and the hydrogen is obtained mainly from natural gas methane. Iron is used as a catalyst. The overall equation for the reaction is: The reaction happens in two stages. This is a good example of the use of transition metal compounds as catalysts because of their ability to change oxidation state. Reactions of iron ions in solution Free iron ions are compeled with water in aqueous solutions. The simplest of these complex ions are: They are both acidic ions, but the iron III ion is more acidic. Reactions of the iron ions with hydroxide ions Hydroxide ions from, say, sodium hydroxide solution remove hydrogen ions from the water ligands attached to the iron ions. When enough hydrogen ions have been removed, you are left with a complex with no charge - a neutral complex. This is insoluble in water and a precipitate is formed. In the iron II case: Iron is very easily oxidized under alkaline conditions. Oxygen in the air oxidizes the iron II hydroxide precipitate to iron III hydroxide especially around the top of the tube. The darkening of the precipitate comes from the same effect. In the iron III case: Reactions of Iron Ions with Ammonia Ammonia can act as both a base and a ligand. In these cases, it simply acts as a base - removing hydrogen ions from the aqua complex. The precipitate again changes color as the iron II hydroxide complex is oxidized by the air to iron III hydroxide. Reactions of the iron ions with carbonate ions There is an important difference here between the behavior of iron II and iron III ions. Iron II ions and carbonate ions You simply get a precipitate of what you can think of as iron II carbonate. If sodium carbonate solution is added to a solution of hexaaquairon III ions, you get exactly the same precipitate as if you added sodium hydroxide solution or ammonia solution. This time, it is the carbonate ions which remove hydrogen ions from the hexaaqua ion and produce the neutral complex. Depending on the proportions of carbonate ions to hexaaqua ions, you will get either hydrogencarbonate ions formed or carbon dioxide gas from the reaction between the hydrogen ions and carbonate ions. The more usually quoted equation shows the formation of carbon dioxide. If you add thiocyanate ions, SCN-, e. Finding the concentration of iron II ions in solution by Redox titration You can find the concentration of iron II ions in solution by titrating with either potassium manganate VII solution or potassium dichromate VI solution. The reactions are done in the presence of dilute sulfuric acid. In either case, you would pipette a known volume of solution containing the iron II ions into a flask, and add a roughly equal volume of dilute sulfuric acid. What happens next depends on whether you are using potassium manganate VII solution or potassium dichromate VI solution. At first, it turns colorless as it reacts. The end point is the first trace of permanent pink in the solution showing a tiny excess of manganate VII ions. The two half-equations for the reaction are: Having got that information, the titration calculations are just like any other ones. Using potassium dichromate VI solution Potassium dichromate VI solution turns green as it reacts with the iron II ions, and there is no way you could possibly detect the color change when you have one drop of excess orange solution in a strongly colored green solution. With potassium dichromate VI solution you have to use a separate indicator, known as a redox indicator. These change color in the presence of an oxidizing agent. There are several such indicators - such as diphenylamine sulfonate. This gives a violet-blue color in the presence of excess potassium dichromate VI solution. The two half-equations are: Once you have established that, the titration calculation is again going to be just like any other one.

Chapter 7 : Iron - Element information, properties and uses | Periodic Table

Full text of "Notes on the chemistry of iron. For professional men, students, iron and steel merchants, and all interested in iron.. For professional men, students, iron and steel merchants, and all interested in iron.

Set up Jar 1 Add a tablespoon of iron filings to the bottom of the jar. Pour enough water into the jar to completely cover the iron filings. This jar acts as your control because it has all the components we commonly associate with rust formation. Do not put on a lid. Knowing that this jar is our control, why would we want to leave the lid off of the jar? Set Up Jar 2 Add a tablespoon of iron filings to the bottom of the jar. Add a teaspoon of calcium chloride to the jar. The purpose of this is to remove all water vapor from the atmosphere. Make sure to screw the jar lid on tightly. Set Up Jar 3 Add a tablespoon of iron filings to the bottom of the jar. What do you think the purpose of adding oil is? Carefully pour water into the jar until a one inch layer is formed. After a couple of seconds, where does the oil layer go? Set Up Jar 4 Add a tablespoon of iron filings to the bottom of the jar. Add enough water to completely cover the iron filings. Add one tablespoon of vinegar. Set all your jars in a quiet place and wait until you see rust in one of your jars. Results You are likely to get results in hours. The filings in Jar 1 and Jar 4 will show rust; the filings in Jar 2 and 3 will not. Jar 4 is likely to have more rust than Jar 1. So how does rust form, exactly? Rust chemistry is fairly straightforward: To get to the oxygen, however, these electrons need to travel through water! Rust appeared on the iron filings in Jar 1 because all reactants were present: The iron was in the filings, the oxygen came from the air, and of course, you added the water. Jar 2 had no water because the calcium chloride removed moisture from the air. Because only oxygen and other gasses in our atmosphere were present in the jar, no rust could be created. In Jar 3, the layer of oil prevented the oxygen in the air from meeting up with the water and iron underneath. In Jar 4, the vinegar created a chemical reaction of its own with the iron filings. This made it easier for the oxygen in the air to react with it and create rust. Disclaimer and Safety Precautions Education. In addition, your access to Education. Warning is hereby given that not all Project Ideas are appropriate for all individuals or in all circumstances. Implementation of any Science Project Idea should be undertaken only in appropriate settings and with appropriate parental or other supervision. Reading and following the safety precautions of all materials used in a project is the sole responsibility of each individual.

Chapter 8 : Notes on the Chemistry of Iron

Iron is easily prone to rusting making its surface rough. Chemically, rust is a hydrated ferric oxide Titanic 's bow exhibiting microbial corrosion damage in the form of 'rusticles'.

Chapter 9 : Full text of "Notes On The Chemistry of Iron"

PREFACE IN pre se nti ng thi s little work to the p ublic, the author has e ndeav ored to embody in plai n la nguage a th or oughly pra cti cal de scripti on of su ch chemi cal meth od s of a naly si s in ir on and steel ma.