

## THE ELECTRIC FURNACE: ITS EVOLUTION, THEORY AND PRACTICE

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Article VI.—Continued.

### OTHER USES OF THE ELECTRIC FURNACE.

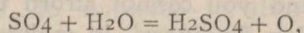
Another method for the treatment of mixed sulphide ores such as those of zinc and lead is the chlorine smelting process of Swinburne and Ashcroft. Although largely a chemical process, the final stage is carried out in an electric furnace, and a short account may, therefore, be given here.

The ore, consisting of sulphides of lead, zinc, iron, and manganese, with some silver, is decomposed by the action of dry chlorine at a temperature of 600° C. or 700° C. in a special vessel called a transformer, forming a fused mixture of chlorides of the metals, while the sulphur comes off in the free state, and can be condensed and saved. The earthy matter, or gangue, from the ore remains suspended in the fused chlorides. Enough heat is produced by the reaction to keep the transformer at the right temperature, which can be regulated by passing the chlorine more or less rapidly. When the transformer is full of chlorides they are tapped out, leaving enough behind to serve as a molten bath into which the ore can be charged, and through which the chlorine can be passed. The molten chlorides are then treated with molten lead, which serves to remove the silver, then with zinc to remove the lead, and the residual chlorides are dissolved in water, separated from the gangue by filtration, the iron and manganese precipitated chemically by the addition of chlorine and zinc oxide, leaving a solution of zinc chloride only. This solution is evaporated, and then fused and electrolysed in a furnace shown in outline in Fig. 19, p. 216, 1906). The products are molten zinc, which is tapped off at intervals, and chlorine, which is compressed and used again for the treatment of fresh quantities of ore. The process is one of great interest, and is applicable to very many complex ores which are difficult to treat by other methods. It is self-contained, and does not require any expensive reagents, as the chlorine for the transformer is produced in the electrolysis of the zinc chloride, but the operations are somewhat complicated, and would need very careful attention. At present the only commercial installation is at a plant of the Castner-Kellner Co., which has a supply of chlorine from other processes, and uses it for the treatment of complex ores as described above, but omits the final electrolysis, obtaining the zinc in the form of chloride. Accounts of this process can be found in the "Electrochemical Industry," Vol. I., p. 412; Vol. II., p. 404; Vol. III., p. 63, and elsewhere.

**Electrolysis.**—The use of a "direct" current for dividing a chemical compound into two component parts has already been mentioned, see "Electrolytic Furnaces," p. 216, 1906, but a few more words may be added here. When a direct or continuous current flows through a fused salt or a solution of a salt in water, the salt or the water is broken up by the current into two parts, one of these being hydrogen, or a metal which is liberated at the cathode or electrode through which the current leaves the liquid, while the remainder of the salt or of the water is liberated at the anode or electrode by which the current enters the liquid. Thus:—

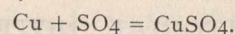
$\text{NaCl}$  (electrolyzed) =  $\text{Na}$  (at cathode) +  $\text{Cl}$  (at anode). That is to say, when fused common salt is electrolyzed, sodium is set free at the cathode and chlorine at the anode.

$\text{CuSO}_4$  (electrolyzed in aqueous solution) =  $\text{Cu}$  (at cathode) +  $\text{SO}_4$  (at anode), or when a solution of copper sulphate in water is electrolyzed copper is set free at the cathode while  $\text{SO}_4$  is liberated at the anode. The final result of the operation will depend, however, upon the nature of the anode. If this is of platinum or carbon, and is not attacked by the  $\text{SO}_4$ , the latter will react with the water of the solution and form sulphuric acid and oxygen, thus:—



and the end products of the electrolysis will be copper at the cathode and oxygen at the anode. If, however, the anode

were made of copper or some other metal that would be acted on by the  $\text{SO}_4$ , this reaction would take place:—



The copper sulphate solution would thus be regenerated, no oxygen would be liberated, and the only result of the operation would be a transfer of copper from the anode to the cathode.

The latter case is exemplified in the electrolytic refining of copper, the anode consisting of impure copper, which constantly dissolves under the action of the current, while pure copper is deposited on the cathode. When it is desired to extract a metal from the fused salt or solution in which it is contained, the anode should, if possible, be insoluble in the solution employed, or if this is impossible it should be inexpensive, as it will be dissolved in proportion as the other metal is recovered.

In the equation given above for the electrolysis of fused common salt, chlorine and sodium are the end products. For the production of sodium an aqueous solution would have been impossible, as the water would react with the sodium, forming caustic soda and hydrogen.

In the electrolysis of a fused mixture of two salts or of a solution of a salt in water, the current will break up the compound which is the least stable; thus in a solution of copper sulphate in water, the current separates the copper sulphate into its components, and not the water; but in a solution of aluminium sulphate it would be the water and not the aluminium salt that would be decomposed. It is necessary, therefore, to employ a solvent that is more stable than the salt it is desired to decompose, or, failing this, to use the pure salt in a state of fusion. This is the reason why the extraction of aluminium from its ore is carried out in a fused mixture of fluorides instead of in an aqueous solution.

In the electrolysis of solutions a definite amount of electricity in passing through the solution will always produce a definite amount of decomposition. This amount is always the same for the same solution, and in different solutions chemically equivalent amounts of decomposition are effected. A current of one ampere flowing through acidulated water for one second will liberate 0.0104 milligrams of hydrogen, and in any other solution the weight of the metal liberated will be 0.0104 milligrams, multiplied by the atomic weight of the metal and divided by the valency of the metal in the particular solution. Thus, the amount of the monovalent metal sodium that would be set free per second would be 0.0104 mg.  $\times$  23, the atomic weight of sodium, or 0.239 mg.; while the weight of copper deposited would be 0.0104 mg.  $\times$  63.2, the atomic weight of copper, or 0.657 mg. in cuprous salts, such as  $\text{Cu}_2\text{Cl}_2$ , in which copper is monovalent, while in the more usual cupric salts, such as  $\text{CuSO}_4$ , in which the metal is divalent, only half that amount would be deposited by the current. The amount of metal actually obtained as the result of electrolysis is frequently less than the calculated weight on account of secondary reactions, such as the metal redissolving in the electrolyte, hydrogen being liberated instead of the metal, leakage of the current, etc.; and the ratio of the metal actually deposited to the theoretical quantity is known as the current efficiency, as it shows what proportion of the current is effective in liberating the metal.

The electrical energy necessary to produce a definite weight of a metal by electrolysis of a chemical compound of the metal depends not only on the number of ampere hours needed to liberate the weight of metal, but also on the voltage of the operation, that is on the electrical pressure needed to drive the electric current through the electrolyte so as to produce the decomposition. Each solution has a definite electrical pressure which must be exceeded before electrolysis will take place, and the working voltage must be decidedly above the minimum in order to drive a