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Chemistry at Dalhousie *circa* 1868

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Chemistry at Dalhousie *circa* 1868

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Abstract: This article describes details of classes at Dalhousie University in 1868–1869, of the life of George Lawson, the first Professor of Chemistry and Mineralogy, and of the wide range of chemical concepts known at that time. A comprehensive set of lecture notes from Lawson's chemistry course, written by a student, Alexander Russell, and held in the Dalhousie University Archives, offers a wonderful insight into the state of chemical knowledge and how it was taught at that time. Lawson began with general chemical principles followed by a detailed discussion of the nonmetals. The second half of the class covered a range of metals followed by a small section on mineralogy and a large section on organic and biological chemistry. Lawson used an older set of atomic masses in which many, but not all, of the elements had masses one-half of the accepted values today. When corrected for these errors, Lawson's formulae, even for complex molecules such as morphine, mostly agreed with contemporary usage. Examples of nomenclature, chemical formulae, preparations, processes and properties are presented. A few examination questions are given also. Even though the concepts involved in understanding chemical structure were just being developed, the breadth and depth of descriptive chemical knowledge at that time was remarkable.

Key Words: chemistry at Dalhousie in 1868, early chemistry in Nova Scotia, George Lawson, organic chemicals from plants, common elements

Introduction

This issue of the Canadian Journal of Chemistry celebrates the 200th anniversary of the founding of Dalhousie University. Contributions from Faculty and Alumni of the Department of Chemistry highlight current aspects of research. Therefore, we felt it was appropriate to recognize the long history of chemistry at Dalhousie by providing a perspective from one hundred and fifty years ago. From 1838 on,¹ chemistry was included in classes on Natural Philosophy while chemistry as a subject has been taught since 1863. At that time, an infusion of cash allowed Dalhousie to reorganize, hire new faculty, and expand its subject offerings. Chemistry was considered important and was included in that expansion with the appointment of George Lawson as the first Professor of Chemistry and Mineralogy. This paper presents a rare view through the eyes of a student of how aspects of chemistry were taught in Canada a century and a half ago.

The 1868–69 Dalhousie Calendar lists only two degrees in the Faculty of Arts,² a general BA degree and an MA degree. The faculty and staff consisted of six professors (including the Principal and Lawson), a tutor for Modern Languages and a janitor. Earning a BA degree required successful completion of four winter terms which ran from late October until early April, with lectures in most classes being held every day of the week. Required subjects included Classics, Mathematics, Rhetoric, Logic and Psychology, Chemistry, Natural Philosophy (Physics), Modern Languages, Metaphysics, Ethics, Political Economy and History, many of which were taught in more than one year. Modern languages could be either German or French. Chemistry was taught in both the second (junior chemistry) and third year (senior chemistry). By way of comparison, at this time McGill University³ offered one Chemistry class of some 45 lectures to first-year BA students. Queen's University⁴ offered one class to third-year BA students, whereas University College, Toronto,⁵ offered several classes to both third- and fourth-year students.

In 1868–69 there were a total of 32 students enrolled at Dalhousie in the four years of the BA degree plus a further 26 “general students” who were permitted to attend classes on payment of the

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requisite fee.² In addition, the 22 students registered in the newly formed Faculty of Medicine were required to take chemistry in their first year of study. The MA degree was awarded to individuals who “maintained a good reputation” for three years after the BA and submitted an acceptable thesis, accompanied by payment for the required diploma fee. According to Department of Chemistry records, Dalhousie awarded the first Master of Arts Degree in Chemistry to James Forrest in 1871 and the initial First Class Honours Degree in Experimental Physics and Chemistry to Ebenezer MacKay in 1886. It is interesting to note that in 1896 Ebenezer MacKay succeeded Lawson as the second Professor of Chemistry at Dalhousie.

Professor George Lawson

Professor George Lawson was an interesting individual. Although formally appointed as Professor of Chemistry and Mineralogy, he was a botanist at heart and published almost all of his scientific work in botany.⁶ Born in Scotland in 1827, Lawson first pursued a career in the law, but gave this up to study science at the University of Edinburgh in 1848.⁷⁻⁹ For the next decade he continued his scientific studies there, acting as a demonstrator in the laboratory of John H. Balfour,⁷ the Chair of Botany at Edinburgh, and working with several scientific organizations, in particular the Botanical Society of Edinburgh and the Royal Society of Edinburgh. At the time of his leaving Edinburgh, Lawson’s work in botany and microscopy was well known and documented in over 50 published articles.⁷

At Edinburgh, the first Chair of Chemistry, the second in Britain after Cambridge, was appointed in 1713, joining botany (1676)¹⁰ in the Medical Faculty.¹¹ From this early beginning, Edinburgh maintained a strong position in chemistry, with such scientists as Joseph Black, discoverer of carbon dioxide, and Alexander Crum Brown, developer of graphical chemical formulae, on the faculty at various times.¹¹ Lawson certainly acquired his knowledge of chemistry during his time at Edinburgh.

It is recorded that Lawson was awarded the D. Phil. degree in 1857 from the University of Giessen.¹² This record also indicates that there was no thesis associated with his degree. A more detailed listing of doctoral degrees with theses and promoters contains no mention of Lawson.¹³ In fact, a listing of personnel at Giessen in 1856–57, “Personal-Bestand der Großherzoglich Hessischen Ludewigs-Universität Giessen”, from the archives shows no record of Lawson. Thus, it is doubtful whether he ever studied in Germany,⁹ and his D. Phil. degree must have been awarded in absentia. This was not uncommon; the majority of the degrees from Giessen in 1857¹² were awarded without submission of a thesis. There is also no record of what published work Lawson submitted as a basis for the awarding of his degree. However, the possession of a degree from Giessen would have helped him to obtain an appointment in chemistry, since at that time, Giessen was acknowledged to have state-of-the-art facilities in chemistry, largely due to the work of Justus von Liebig, who was widely recognized for his teaching and broad-based research in chemistry.

In 1858 Lawson accepted an appointment as Professor of Chemistry and Natural History at Queen's College in Kingston, Upper Canada. Over five years in Kingston, he taught chemistry and set up laboratories in chemistry and botany. Many students were influenced to take up botany through contact with Prof. Lawson,⁸ who was also instrumental in establishing the Botanical Society of Canada.¹⁴ Indeed, he is considered by many to be the “Father of Botany in Canada”.

From 1863 until his death in 1895, Lawson taught chemistry in both the Faculty of Arts and the Medical Faculty at Dalhousie. A practical man, he became involved with agriculture. He not only owned and ran a model farm a short distance from Halifax, but also acted as secretary to the Central Board of Agriculture of Nova Scotia and subsequently as Secretary of Agriculture for the Province. He was also a member of many learned societies, including the Royal Society of Canada, of which he was a charter member and president in 1887–88,⁷ and the Nova Scotian Institute of Science, of which he was president at his death.⁸ In addition to these teaching and other duties, he continued to publish scientific research, mainly in botany. All told he authored over one hundred publications,

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which included chemical topics such as nomenclature, water analysis, chemistry and heat, gold processing and the discovery of a dye from the black spruce Aphis.⁶

Changing chemical concepts in the eighteen sixties

To provide some perspective for the chemistry in Lawson's lectures, which preceded the famous periodic table published by Dmitri Mendeleev in 1869,¹⁵ it might be helpful to list a few developments in the state of chemical knowledge around that time. In 1852, Edward Frankland is credited with proposing the theory of chemical valency,¹⁶ and the tetravalency of carbon was suggested by August Kekulé in 1857–58.¹⁷ In the 1860s and beyond, the Theory of Structure, which recognized and exploited specific atomic connectivity, was developing,¹⁸ and the ring and bonding structure of benzene was suggested by Kekulé in 1865.¹⁹ In the mid-1850s Kekulé and other notable German chemists published using the old atomic mass scale championed by Berzelius in which C = 6 and O = 8.¹⁷ At the Karlsruhe Congress of 1860, held as an attempt to bring some order and consistency to the state of chemistry at that time, Stanislao Cannizzaro used Amedeo Avogadro's ideas about diatomic elemental gases and gas volumes to propose a consistent set of atomic masses, which included C = 12 and O = 16.^{20,21} Although not immediately accepted, these masses eventually became universally employed.

Chemistry as taught by Lawson in 1868–69

A detailed account of the 89 lectures presented daily by Prof. Lawson was recorded by Alexander Russell,²² one of nineteen BA students in the chemistry class who later became a Minister in the Presbyterian Church. Russell's notebook,²³ located in the Dalhousie University Archives, describes a remarkably comprehensive presentation of chemistry, with frequent lecture demonstrations. The lecture notes covered 185 pages of fine handwriting. As in any note-taking situation, it is possible that the notes may include some misrepresentations of the material

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presented. Also, it must be noted that some of the many misunderstandings prevailing at that time, in particular, the problem of the atomic masses, explained below, were reflected in the recorded material. The recommended text for junior students was "Practical Chemistry" by Stevenson Macadam (W. & R. Chambers, Edinburgh); for senior students, "Fownes Chemistry for Students" by George Fownes, edited by Robert Bridges (Blanchard & Lea, Philadelphia) was recommended. What follows is a selected summary of the large amount of material detailed in Russell's notes.

Names and Formulae

The elements were described using the same symbols used today but atomic masses were called "equivalent numbers". While some of the atomic masses were very close to those accepted today, a considerable number were half of current values (e.g. C = 6 and O = 8) and a few were inexplicably different (see Table 1). This led to differences in chemical formulae compared to present practice. If a compound contained only elements with correct atomic masses or masses incorrect by a factor of two, the chemical formula would agree with the contemporary formula. If a compound contained a mixture of these types of elements, the number of atoms of some elements in the formula would be wrong by a factor of two. For example, Lawson presented the formula of water as HO, not H₂O. The formula HO gave the correct weight percent of the elements in water assuming atomic masses of H = 1 and O = 8.

In addition to the classification of elements as metals or nonmetals, some elements with similar properties were grouped, such as Na, K, Mg, Ca, the "alkali class". Compounds such as oxides and sulphides were recognized, and formulae were written as today using subscripts to indicate stoichiometry. Basic oxides were called "oxides", while acidic ones were called "acids", but not all acids contained oxygen, e.g. HCl. When writing formulae for binary compounds, the metal symbol was written before the nonmetal, whereas in oxides, the basic oxide was first and the acidic one second. For example calcium sulphate could be written either as CaSO₄ or as CaO,SO₃, but potassium sulphate was written as KSO₄ or KO,SO₃. These formulae reflect the correct mass for

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potassium and the half values of the masses for calcium, oxygen and sulphur. Formulae for more complex molecules (called “bodies” by Lawson) were known, such as morphine, $C_{34}H_{19}NO_6$; when corrected for the erroneous atomic weights of carbon and oxygen, the formula is $C_{17}H_{19}NO_3$, as is accepted today.

Stoichiometry, thermochemistry and electrochemistry

Using combustion as an example, Lawson introduced fundamental principles. In particular, he told the students that matter could not be created or destroyed and that one type of matter reacted with another in fixed proportions. For example, he explained how 1 pound of carbon would react with $2\frac{2}{3}$ pounds of oxygen to produce $3\frac{2}{3}$ pounds of carbon dioxide, plus enough heat to convert 122 pounds of water at 60° F to steam at 212° F, which could be used to do work. The stoichiometry of the reaction is in agreement with values used today, while the thermochemistry is off by less than 6%. There is no record in the notes as to how the numbers were derived, but presumably they were experimental values from other work. While the units used by Lawson are antiquated, he realized that there were two temperature scales in use, Centigrade (now Celsius) and Fahrenheit, and that measurement of temperature depended on the expansion of materials, such as a metal rod or a column of mercury.

After noting that combustion of coal in air produced carbon dioxide, and combustion of carbon compounds containing hydrogen also produced water (shown as HO), he pointed out that all living animals respire, consuming oxygen and producing carbon dioxide. To maintain balance on the planet, he noted that plants absorb carbon dioxide, release oxygen and convert carbon dioxide into coal over a long period of time, thus storing energy for later use. It is sobering to realize that this was taught to students 150 years ago and is of equal importance in the world today; sadly however, this is unrecognized by many.

Lawson also indicated that combustion produced both heat and light. Within this context, he pointed out how various forces could be converted into other forces, such as steam generation

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(above), or the use of electricity to overcome chemical affinity in the electroplating of copper from a solution of cupric sulphate. Other topics discussed included batteries made from zinc and copper, and the flow of electricity in a coil used to produce a magnetic field.

Properties and reactivity of common nonmetallic elements

Oxygen

After briefly mentioning the discovery of oxygen by Priestley, Lawson pointed out that it could be obtained by applying a force (i.e. heat) to many oxygen-containing substances, although it was usual to do this with chlorate of potassium, which he wrote as KO, ClO_5 . Lawson demonstrated the preparation of oxygen and did the usual experiment in which a glowing wood splint placed in an atmosphere of oxygen would burst into flame. He warned that saltpetre (potassium nitrate) could behave similarly, but was dangerous on contact with combustible materials and he classified materials as either combustible or not. He suggested that oxygen had powerful attraction for other elements, reacting with all except fluorine, although that perhaps it did, but this had not been discovered. Fluorine was dangerous to handle.

Lawson explained how certain metals could react with oxygen in the air, but the oxides formed then inhibited further reaction; this was not true for iron, which did react further. Oxides could be basic, neutral or acidic, and the more oxygen the latter possessed, the more acidic they became. Finally, Lawson discussed the dilemma of oxygen and ozone. It was known that these were different, but the difference was ascribed to them being allotropes, such as red and white phosphorus. Ozone was formed after lightning strikes, was a powerful oxidant and, like chlorine, an excellent disinfectant and bleach. Ozone was easily detected by starch paper soaked with iodide of potash and the depth of the resulting blue colour indicated the concentration.

Hydrogen

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In a discussion of hydrogen, the lightest element with an equivalent number of 1, Lawson noted that it combined with other elements in the proportion of the relative equivalent numbers. Hydrogen and oxygen could be generated at the different poles in the system by the electrolysis of water. A different way was to add a small piece of sodium metal to water. He also pointed out that other less reactive metals, such as zinc, would do the same when placed in a dilute acid and he demonstrated both effects. He warned that a mixture of oxygen and hydrogen was perfectly stable if undisturbed, but would explode if subjected to any forces such as an electric spark, light, sudden compression or finely divided platinum. In another of many in-class demonstrations, Lawson exploded a mixture of hydrogen and oxygen. Hydrogen and oxygen could be ignited in a controlled jet to produce an intensely hot flame. If the flame was confined in a hollow glass tube, sound would be generated at a frequency which could be tuned by the size of the tube. Lawson explained this as the flame being a continuous series of small explosions, thus generating sound. Burning of hydrogen produced its oxide, water. Hydrogen was the lightest known gas with 100 cubic inches weighing only 2.14 grains and in principle it could float on top of other gases, but in fact mixed with them because of diffusion. He mentioned that hydrogen had a second oxide, HO_2 , which was an excellent bleaching agent. Hydrogen could form an oxide which was both acidic or basic.

Nitrogen

Lawson told the class that the properties of nitrogen were best characterized by negatives – it did not support either life or combustion, was colourless, tasteless, non-toxic and insoluble in water. He divided the components of air into major, nitrogen and oxygen, and minor, carbon dioxide, water vapour, hydrogen sulphide, nitric acid, ammonia, ozone and sulphur dioxide. Nitrogen had a weak affinity for oxygen, but it did form five oxides, each with an increasing proportion of oxygen. These were given as NO (nitrous oxide, laughing gas), NO_2 (nitric oxide), NO_3 (hyponitrous acid), NO_4 (nitrous acid) and NO_5 (nitric acid); these formulae have a correct atomic mass for nitrogen, but an incorrect one for oxygen.

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Heating nitrate of ammonia to 300° F in a retort gave water and nitrous oxide, because hydrogen had a greater affinity for oxygen than did nitrogen. Humphrey Davy had found that the gas could cause a form of intoxication, hence the name laughing gas. Nitrous oxide acted on the human muscular system, not the nervous system, unlike the actions of morphia or strychnine. Nitric oxide was prepared by heating nitrate of copper and was a gas with a strong disagreeable odour. It could be oxidized to NO_3 or NO_4 , which were unimportant.

Nitric acid was a strong oxidizing agent and could be formed by heating 2 parts of nitrate of soda with one part of sulphuric acid; nitrate salts were common in the deserts of South America. Nitrogen and oxygen do not combine directly, only in an electrical discharge. In large cities with little pure air and much filth, nitric acid can even be found on the walls of buildings. Nitrates are formed by decaying organisms and have been found in wells close to churchyards, a point relevant to environmental concerns today. Nitric acid was normally obtained as a 70% aqueous solution, but it can also be prepared in essentially anhydrous form. It is colourless when pure, but is brown when contaminated with nitrogen dioxide. It reacts vigorously with most metals and with living or dead tissues.

Carbon

Carbon was interesting as it is found in several forms, wood charcoal, lamp black, animal charcoal, coke, graphite (soft and considered to be the normal form of carbon) and diamonds (the purest form of carbon and very hard). Animal charcoal had the greatest porosity and was used to purify sugar in its manufacture. Lawson presented a table showing the relative amounts of gases that could be adsorbed by animal charcoal, with ammonia being the greatest and hydrogen being the least.

Carbon was known to form two oxides, CO (carbonic oxide) and CO_2 (carbonic acid). The former was formed during incomplete combustion of charcoal and was poisonous to the lungs, but it could also be made by treating ferro cyanide of potassium with sulphuric acid. Lawson explained

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that this reaction was complex, but there was only enough oxygen present to form CO and not CO₂. The latter was a heavy gas, formerly known as “fixed air” because it could be fixed as solid calcium carbonate when bubbled into lime water. It could be released from this, or from marble, by heating or by treating the solid with acids such as sulphuric or hydrochloric. It was freely soluble in water to form the weakest known acid. Carbon dioxide was also produced by fermentation, respiration, volcanic eruptions and in mineral springs.

Lawson moved on to discuss compounds of carbon with hydrogen, while pointing out that most of these would be discussed as organic chemistry. He focused on the two simplest hydrocarbons, CH₂ (light carburetted hydrogen) and C₄H₄ (olefiant gas). The former had a long history of use for illumination and was made commercially by dry distillation of bituminous coal. In a detailed discussion he described how impurities such as ammonia, hydrogen sulphide, sulphur trioxide and carbon dioxide were removed and analyzed to ensure that the coal gas was pure. Lawson explained how light came from combustion of carbon and not from combustion of hydrogen.

The halogens

The elements chlorine, bromine, iodine and fluorine form a natural grouping. In nature none is found as a free element, but always as a salt, the first three in salt deposits and sea water, the last in the mineral fluorspar. Of these, chlorine was covered most extensively initially by pointing out that the action of a strong acid on the muriate of soda produced hydrogen chloride. To make chlorine, the hydrogen needed to be removed, which was done by adding manganese dioxide. Chlorine had a great affinity for all elements and was dangerous to handle; it attacked the lungs causing violent bursts of coughing and usually permanent damage. Chlorine could be liquefied under pressure, but it did not congeal even at the coldest temperatures. Chlorine reacted unusually with some organic compounds such as turpentine; one hydrogen atom was removed and replaced by a chlorine atom. We now know that this is an example of free radical halogenation at an allylic centre. Chlorine was soluble in water producing unstable hypochlorous acid, but the sodium salt of this acid was stable

and, like chlorine, was an excellent bleach and disinfectant. Lawson described four oxides of chlorine, as well as their associated acids, each with an increasing oxygen content.

Hydrogen chloride was of great importance, four volumes of which were formed explosively when two volumes each of chlorine and hydrogen were exposed to sunlight. It had a great attraction for water, which could absorb 480 times its bulk of hydrogen chloride. Hydrochloric acid would dissolve most metals except gold and platinum, but if mixed with nitric acid to make aqua regia, even these could be dissolved. It was chlorine that was set free in this acid which attacked these noble metals.

Bromine could be released from its sodium salt by the action of chlorine and was easily separated because of its solubility in "sulphuric ether" (ethyl ether). Where chlorine was a green gas, bromine was a brown liquid and iodine a violet solid. In general bromides were much more soluble in water than chlorides. Iodine was unusual because on heating the solid converted directly to the vapour. It was found in many marine organisms, especially kelp.

Fluorine was different from the other halogens. Only hydrogen fluoride was discussed. Lawson prepared this from calcium fluoride and sulphuric acid to demonstrate how it etched glass. He emphasized that hydrogen fluoride was dangerous and caused severe burns to human tissues.

Sulphur, phosphorus, silicon and boron

Sulphur was discussed at length, initially by describing its allotropes, crystalline, vitreous (glassy) and amorphous. The first was soluble in bisulphide of carbon, the second was "elastic" and the third was "earthy". Sulphur was found in volcanoes, pyrites and in all living matter, especially albuminous matter such as eggs. Lawson described seven oxides of sulphur, with emphasis on the dioxide, which was easily oxidized to the trioxide. Sulphur dioxide and its aqueous solution were good bleaching agents, but differed from the bleaching action of chlorine. Sulphuric acid was the strongest known acid, formed salts with alkalis and oxidized metals to form salts, as $\text{Zn} + \text{H}_2\text{SO}_4$ gave ZnO, SO_3 . Hydrogen sulphide, written as H_2S , was found in decaying vegetable or animal matter

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and was discussed in detail. It was used as a test for metals, with many of them giving insoluble, coloured sulphides. Lawson made only a brief mention of selenium, pointing out that it was always found in conjunction with sulphur and that "HSe smelled even worse than HS".

Phosphorus never occurred in nature in the free elemental form, but was always combined with oxygen, although occasionally it was found as a phosphide. Phosphorus mostly occurred as phosphate of lime, $3\text{CaO}, \text{PO}_5$, which was found in soils and bones. The element was essential for life and even occurred in the brain. Lawson described the white and red allotropes of phosphorus. The former was toxic, smelly and soluble in bisulphide of carbon; it burned in air and glowed in the dark. This latter property was a sensitive test for the material. By contrast, red phosphorus was more stable and much less toxic or soluble. The element burned with a brilliant flame and emitted copious quantities of white smoke. It was used in matches. Lawson described four oxides of phosphorus including the most common, " PO_5 , phosphoric acid". Next to nitrates, soluble phosphates were the most important components of manures for growing plants.

Finally Lawson briefly discussed silicon and boron. The former occurred mainly as the dioxide, the only known oxide. Glass was a silicate of potassium. Silicon also formed a compound with fluorine, which he wrote as " SiF ". Boron was somewhat like silicon and was found in nature as borates (salts of boracic acid). These could be found in volcanic springs in places like Persia. Borates have been used to render metals fusible or as "a bromide flux". These two elements completed coverage of the nonmetals; the lecture on December 23, 1868 was the last of the calendar year.

Properties and reactivity of selected metallic elements

Lawson spent the first four weeks of the new year discussing a total of thirty-two metallic elements. He first covered the alkali and alkaline earth metals as groups, followed by most of the first row of transition metals in reverse order to their atomic masses, then a mixture of elements in no particular order, and finally some of the noble metals. If judged by the volume of notes, Lawson

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may have considered Au, K, Al, and As to be the most important metals and Ti, V, Li, Mo, W, and Pd to be the least important. It is more likely that these two groups represented the metals about which the most and least were known. All this was presented in a descriptive manner and included such topics as properties, methods of preparation, uses, chemical reactions, and analytical tests. Four examples of elements important to Nova Scotia are included below.

Sodium

Lawson noted that sodium metal behaved in a manner similar to potassium. It was soft, lustrous when first cut, soon reacted with air and burned with a yellow flame when placed in water, resulting in an alkaline solution. Two oxides of sodium were mentioned: NaO, caustic soda – by far the more important of the two, and Na₂O. Various methods of preparing carbonate, bicarbonate (important in baking), and sulphate of soda were described. Originally sodium salts were obtained from plant ash but more recently sea salt (NaCl) had replaced this source.

Iron

Lawson described iron as a very important element with a wide distribution in various rocks and soils as well as in our bodies, especially in blood. Iron ores included both FeO and Fe₂O₃, “bog ore”, and iron pyrite (FeS₂). He suggested that magnetic iron ore (magnetite) was a mixture of FeO and Fe₂O₃. Iron could be obtained by roasting some of the iron ores in air to form iron oxide. The oxide was then placed in a blast furnace along with coal and a stream of air passed through the furnace, a structure described as 50 feet in height and 17 feet in diameter. The coal provided heat and took the oxygen away from the iron, providing elemental liquid iron, which was removed from the furnace.

Since iron was such an important element, tests were needed for its detection. Two tests were described. The first involved adding the ferro cyanide of potassium to an iron-containing solution, which resulted in a blue colour and a dark blue precipitate. The second used sulfo cyanide

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of potassium (potassium thiocyanate) as a reagent to form a deep red solution indicating the presence of iron. This was described as a very sensitive test, useful for mineral waters. Other aspects of iron chemistry were to be discussed in later sections of the course, where appropriate.

Arsenic

This element had been known for a long time. Arsenic shared some properties with phosphorus, was quite poisonous but otherwise had a high metallic lustre and electrical conductivity. It was said to burn in air with a bluish flame and its vapours smelled like garlic. The main source of the element was mispickel, a mixture of As, Fe, and S (today known as arsenopyrite). Other sources included the arsenides of Ni or Co. These ores could be roasted in air in a “reverberating furnace” to give AsO_3 . Solid arsenic was collected on the walls of long, cool tubes associated with the roasting.

Two oxides of arsenic were known: AsO_3 (arsenous acid), the more important, and AsO_5 (arsenic acid). Arsenous acid was not very soluble, even in boiling water. Thus, even though arsenic was recognized as poisonous, if this property were of interest, strychnine was recommended over AsO_3 for solubility reasons. To prepare soluble arsenic, salts like KO,AsO_3 should be employed – this was the form of arsenic used in medicine. A mixture of As and Cu (probably copper arsenite) produced a green pigment used in paints and in colouring women's dresses.

An interesting flame test for arsenic in a solid sample was described. The sample was placed in a glass tube through which hydrogen was passed, generated by adding zinc to sulphuric acid. This was lit at the end of the tube. The sample was then heated with a “spirit lamp”. If it contained arsenic the flame changed to a “peach-blossom” colour and white AsO_3 was given off. Holding a porcelain cup near the flame resulted in a coating of solid As, the “Arsenical Mirror”. This is known as the Marsh test and depends on the formation of arsine in the heated tube. Another test involved placing a copper foil in a solution containing arsenic. Hydrochloric acid was added to the sample, which was heated carefully, and a dark coating of arsenic would form on the copper. The copper foil

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was then transferred to a dry test tube and heated to drive off most of the arsenic. If a red powder remained, described as an alloy of copper and arsenic, the presence of arsenic was indicated.

Gold

Gold was often found in the elemental form, known by its bright colour, and sometimes found in combination with As, Sb, Cu, Os, or Ir. In Nova Scotia it was found in quartz veins. Lawson went into detail about how gold might be formed in such veins. He described the mercury amalgam method of removing gold from its ores. Solid gold was retrieved from the amalgam by volatilizing the mercury. He did not mention the health hazards of such a procedure. Gold was characterized as a yellow coloured metal that never tarnishes, extremely ductile and malleable, and that melts at 2076° F (contemporary value = 1947° F). It was a good conductor of heat and electricity.

Two oxides of gold were mentioned, AuO and AuO₃. It was also noted that gold could form chloride, bromide, and iodide salts. It did not dissolve in acids except for aqua regia. This formed the basis of the “acid test”, in which a drop of nitric acid (NO₅) was added to the metal. If it dissolved, it was not gold. Another test involved adding sulphate of iron (FeO,SO₃ – ferrous sulphate) to a solution. If gold was present, a metallic precipitate would form. Other tests for gold in nature were described. Iron pyrite (“fool’s gold”) could be differentiated from gold using a hammer – the iron pyrite would fracture, gold would not. Or a drop of nitric acid would form fumes with iron pyrite but not with gold. Copper pyrites could also appear to be gold – these fused with heat while gold did not (unless quite hot). Finally Lawson mentioned that mica could be mistaken for gold due to its glitter in rocks but did not propose a test to determine its presence.

Mineralogy

Five lectures were given to the discussion of mineralogy. Lawson believed that the importance of mineralogy had waned somewhat over the fifty-year period preceding his lectures. He

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made a distinction between minerals, which have a fixed composition, and rocks, which consist of mixtures of minerals. He suggested that geologists were more interested in rocks than minerals.

Minerals were classified into seven orders. Each order consisted of various families. Lawson described the properties of the major minerals in the more important mineral families. These properties consisted mainly of cleavage and fracture, hardness, colour and lustre, and chemical composition. He recognized that the colour of a mineral often depended on the impurities present in the mineral. As an example, orthoclase belonged to the felspar (Lawson's spelling) family of order one, oxidized stones. It was said to be a double silicate of aluminum and potassium. Lawson did not list the exact composition, known today to be KAlSi_3O_8 , but indicated that it consisted of alumina, silica, and potassium oxide. The other properties consisted of specific gravity, 2.5; hardness, 6, (somewhat softer than quartz which had a hardness of 7 to 7.5 on the scale used by Lawson); and a colour ranging from colourless to white, but light yellow or pinkish due to impurities.

The material presented on mineralogy was entirely descriptive.

Organic and biological chemistry

Lawson devoted almost one third of his lectures to this theme, paying particular attention to biological concepts, not surprisingly since most organic compounds known at that time had a biological origin. Since he was teaching medical students, he devoted time to anatomical concepts, as well as to therapeutic uses and toxicity of compounds. He distinguished between tissues and products. Mostly the elements involved were C, H, N, O, P and S.

His introductory remarks were especially interesting: "The object of the chemist is to bring together the elements. Each of the elements we find has a certain amount of 'force'. The chemist has to bring them together and they will form a new body. In order for particles to combine they must be in contact, e.g. if metals are mixed and heated they will form alloys. Bodies can be broken up by disturbing the chemical affinity, by application of some other force, such as heat, light, electricity or mechanical action. Each element then takes back the chemical affinity that it originally possessed. In

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each case, chemists are only arranging bodies, making possible combinations, not by direct force not by the Alchemists Philosopher's Stone, but simply by placing them together. In inorganic chemistry we always know what will come, what can be done in nature can always be done by the chemist. In organic chemistry the same compounds are never formed from their elements, they are formed differently. Compounds are formed because there is something more powerful than chemical affinity, maybe the vital force."

Clearly at that time, the concepts of chemical structure were in flux and Lawson was obviously unsure of how organic compounds and structure were related. His suggestion of "vital force" appears to be in relation to the fact that plants and animals can make different organic compounds and today we know that this is mostly done by enzyme-catalyzed reactions. He proceeded to give an example of how plants could take carbon dioxide and water from the air and by "some force" could split the carbon dioxide to give oxygen with "CHO" left behind, but this was not further deoxygenated. Different plants had different powers of making secretions, e.g. balsam firs made oils containing only carbon and hydrogen, but plants could also make compounds consisting of carbon, hydrogen and oxygen.

Lawson described the general anatomical details of both plants and animals. Plants were composed of cells which could be seen under a microscope. Cells could contain sugar, starch and lignin, while leaves contained chlorophyll. Cells have a nucleus which enables new cells to be formed. Different tissues could be classed as cellulosic or woody, whereas animals had different classes of tissues. Alveolar tissues formed the basis of animal structure, adipose tissues contained fats, cartilaginous tissues contained hyaline substance, while osseous tissues were filled with CaO , CO_2 and 3CaO , PO_5 . All tissues were permeated with liquids.

The saccharine group

After this general introduction, Lawson moved on to discuss the saccharine group, including starch, lignin, cane sugar, fruit sugar and milk sugar, along with their empirical formulae. Starch,

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found only in plants, was important and he gave details of its manufacture and detection by chemical tests. He pointed out that pure sugar did not ferment, but in the presence of yeasts or fungi, sugars would ferment to alcohol, carbon dioxide and water; the alcohol could be obtained by distillation. This provided an entry to the homologous series of alcohols, which concluded with “melisic alcohol, $C_{60}H_{122}O_2$ ”, today known as 1-triacontanol. The next concepts defined the same homologous series of radicals, written as C_nH_{n+1} , beginning with methyl, then ethyl, followed by derivatives of these such as alcohols, aldehydes and acids. Lawson introduced the radical formyl as C_2H , which he pointed out could be considered as derived from methyl by loss of hydrogen. The simplest acid, formic acid, was found in red ants. Related to this, he mentioned chloroform, which produced a state of insensitivity to pain. Ether does this also, but is less powerful.

Discussion continued to “acetyl” and higher homologs; acetic acid was written as “ $C_4H_3O_3 + HO$ ”. Oxalic acid could be formed by oxidizing “amyl” or fermenting potato or grain starch. One hundred pounds of starch could be fermented from which 150 pounds of crystalline oxalic acid could be obtained. This acid was toxic with the antidote being lime or chalk which formed an insoluble salt. Oxalic acid was good for removing ink stains, it did not grow mould. Lawson also discussed other plant-derived compounds, including the conversion of wood fibres into paper.

Fats, oils and organic acids

A description of fats and oils, such as stearine, oleine and palmitine, led to drying oils (poppy, linseed, cod liver, hemp, etc.), non-drying oils (olive, almond, rape, animal, etc.) and intermediate oils such as ricine (castor oil). Most oils had an even number of carbon atoms, although margarine yielded margaric acid which had an odd number. The drying oils absorbed oxygen from the air to make resins, with linseed oil being the best at this. By contrast lubricating oils were unreactive and did not absorb oxygen.

Fats could be saponified by heating with alkali such as sodium or potassium hydroxide to give the salt of the fatty acid. Potassium salts were more soluble and made soft soaps, whereas

sodium salts gave hard soaps and calcium salts were insoluble in water. The topic of oils was concluded by discussing volatile oils, such as balsam, camphor, cedar, coriander, ginger and lemon; all were plant-derived.

Lawson moved on to discuss polybasic organic acids, those which combined with two or more equivalents of base; among others, tartaric, malic and citric acids were featured. He described a range of derivatives of tartaric acid, including among others, cream of tartar, tartar emetic and Rochelle salt. Lactic acid was obtained from fermenting milk, which contained lactose and casein, the latter being important for building muscle. Indeed, milk was the perfect food. Meconic acid was found in the milky juice of certain poppies, from which opium was obtained. Tannic acid was found in the bark of certain trees and was used for tanning leather, after which Lawson discussed glues and inks.

Nitrogen-containing organics

Alkaloids were described as nitrogen-containing, strongly basic poisons found in plants. They could be classified as either volatile oils, or crystalline solids. Examples were given of both types. Nicotine, a component of tobacco was important and for customs and excise purposes, tobacco leaves could be identified by treating with boiling water and smelling the extract. Lawson described the extraction of nicotine from tobacco leaves and pointed out the toxicity of the alkaloid; even a small quantity was able to kill a large animal. He discussed some fifteen alkaloids in five lectures, concluding with caffeine and theobromine. He then discussed organic dyes followed by the chemistry of wood and its derivatives such as gun cotton.

Coverage then turned to other nitrogen-containing organics, the first being cyanogen, which Lawson also called "bicarbide of nitrogen", a misnomer, since in chemistry today carbides have carbon combined with a less-electronegative element. He explained that cyanogen was important in providing support for the theory of radicals, although the theory of radicals was supplanted some ten years earlier by the theory of valency propounded by Frankland and Kekulé. Heating cyanogen with

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potassium gave potassium cyanide, one of a range of compounds involving “cyanide”, such as potassium ferricyanide and Prussian Blue. Lawson also described the isolation and properties of hydrocyanic acid, a weak acid which just turned blue litmus red and was the most toxic substance known.

Medical connections

Lawson did not mention the word “biological”, but devoted the last eight lectures of the class to biological and medical aspects of animals and how these depended on plants for survival. He pointed out that animal matter gave a strong smell of ammonia when burned, but plant matter did not. Albuminoids included proteins, albumin, fibrin (muscle tissue), casein and legumin and were plastic materials. Gelatinous substances included gelatin, chondrin, osseine, sclerotin and cartilage. He described functions for most of these materials along with properties, such as pure albumin could be obtained from egg white and was precipitated by the addition of acid. Lawson presented a table classifying CHO-containing organic molecules into seven groups. All of these, whether animal or plant-derived, came from carbon dioxide and water, but also with a deoxygenation step that only plants could do. For example, oxalic acid $C_4H_2O_8$, was derived from $4CO_2 + 2H_2O$, but two oxygen atoms needed to be lost.

He reiterated how animals get food from plants, which are the “servants of animals” while plants generate “respiratory food” by means of the sun and air. He discussed how food needs to be well masticated by saliva and then continued to discuss digestion, the stomach, intestines and pancreatic fluid. Tape worms absorb nutrients, but have no stomach. The final topics were urea and cyanic acid together with blood, which contained iron. Identification of blood stains was normally done using a microscope.

Final Examinations

The final examinations for all Dalhousie courses were printed in the Calendar of the following year.² Students in Professor Lawson's class of 1868–69 had access to the junior and senior examinations in the previous year. A period of four hours was provided for these written examinations, which consisted of eight to twelve questions. In general, one question on each examination involved a mathematical calculation.

From these examinations it is obvious that the junior and senior students were responsible for different parts of the material. Russell and his fellow students were expected to have memorized the equivalent numbers of all the common elements as well as all the descriptive chemistry that Lawson had covered. The Junior Examinations mainly covered inorganic topics and included questions on nomenclature, definitions, chemical formulae, as well as descriptions of chemical processes and preparations. For example, a question on the Junior Chemistry Examination of 1868 reads "Describe fully the Chemical changes by which Common Salt is converted into Carbonate of Soda, in the ordinary methods of the Soda Manufacture." In 1869, Russell was asked about silver in solution, how and in what form it could be precipitated. The final part of this question was to provide a calculation as to how one could estimate the amount of metallic silver.

Senior Examinations emphasized organic and biological chemistry with greater weight placed on description of more complex processes. The Senior Examination in 1868 required a calculation as follows: "Found grains 3.05 Tersulphide of Antimony; suppose the whole of the Sb to have originally existed in the form of $\text{KO}, \text{SbO}_3, \text{C}_4\text{H}_4\text{O}_{10} + 2 \text{HO}$, how much Tartar Emetic would the above 3.05 gr. represent?"

Concluding remarks

The remarkable feature of Lawson's lectures is the level of detail that students were expected to learn. We were surprised by the extensive level of knowledge at that time. However, not all the material was fully up to date. Keeping abreast of new developments was undoubtedly

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challenging since few chemical journals were available at that time and the availability of such material in Halifax is unknown. During the lectures, Lawson mentioned other texts by Wilson (Chemistry) and Graham (Inorganic Chemistry). Students could sign up for an extra course in practical chemistry (3rd year) which used C. R. Fresenius' classic text on Qualitative and Quantitative Analysis (Vieweg, Braunschweig). It was clear that there was confusion in understanding some chemical concepts. Very little was said about chemicals from coal while petroleum-based chemicals were yet to be developed. The lectures were striking in how Lawson presented the natural balance between plants and animals and how virtually all of the organic chemicals were plant-derived, concepts to which much attention is now being focused, 150 years later.

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Table 1. Lawson's Table of the Elements from Russell's 1868–69 Notes

Element ^a	Lawson's Symbol	Eq. No. ^b	Atomic Mass ^c	Element ^a	Lawson's Symbol	Eq. No. ^b	Atomic Mass ^c
Hydrogen	H	1.0	1.0	Arsenic	As	75.0	74.9
Lithium	Li	6.5	6.9	Selenium ^d	Se	39.5	79.0
Beryllium ^e	Be	6.9	9.0	Bromine	Br	80.0	79.9
Boron	B	10.9	10.8	Rubidium	Rb	Unk	85.5
Carbon ^d	C	6.0	12.0	Yttrium	Y	Unk	88.9
Nitrogen	N	14.0	14.0	Zirconium ^e	Zr	33.6	91.2
Oxygen ^d	O	8.0	16.0	Silver	Ag	108.0	107.9
Fluorine	F	19.0	19.0	Cadmium ^d	Cd	56.0	112.4
Sodium	Na	23.0	23.0	Tin ^d	Sn	58.8	118.7
Magnesium ^d	Mg	12.0	24.3	Antimony	Sb	120.3	121.8
Aluminum ^d	Al	13.7	27.0	Iodine	I	127.0	126.9
Silicon ^e	Si	21.3	28.1	Cesium	Cs	Unk	132.9
Phosphorus	P	32.0	31.0	Barium ^d	Ba	68.6	137.3
Sulphur ^d	S	16.0	32.1	Lanthanum	La	Unk	138.9
Chlorine	Cl	35.5	35.5	Cerium	Ce	Unk	140.1
Potassium	K	39.0	39.1	Didymium ^f	D	48.0	---
Calcium ^d	Ca	20.0	40.1	Terbium	Tb	Unk	158.9
Chromium ^d	Cr	26.7	52.0	Erbium	E	Unk	167.3
Manganese ^d	Mn	27.6	54.9	Gold	Au	197.0	197.0
Iron ^d	Fe	28.0	55.9	Mercury ^d	Hg	100.0	200.6
Cobalt ^d	Co	29.5	58.9	Thallium	Thl	Unk	204.4
Nickel ^d	Ni	29.6	58.7	Lead ^d	Pb	103.7	207.2
Copper ^d	Cu	31.7	63.5	Thorium ^e	Th	59.6	232.0
Zinc ^d	Zn	32.6	65.4	Uranium	U	Unk	238.0

^a Ordered by atomic number^b Lawson's equivalent to atomic mass^c Present values^d An element with the atomic mass divided by two^e An element with an incorrect atomic mass^f Didymium, actually a mixture of praseodymium and neodymium, was considered an element at that time.

Chemistry at Dalhousie College and University, 1868–69

Graphical Abstract

34x18mm (300 x 300 DPI)

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