

## The fabulous history of the solvated electron

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**ABSTRACT** The history of sciences is rich in founding experiments, which have given rise to such developments that they constituted full-fledged disciplines. Much more rarely, knowledge coming from very distant horizons converged at a moment to offer a very new vision of old results and to provoke an explosion of new research. This was the case for the solvated electron. But where does such a concept come from, of an electron isolated in a liquid whose molecules are polarized by its charge?

**KEYWORDS** Solvated electron, Properties, Structure models, Metal-ammonia solutions, Radiation chemistry

**CLC** TL13

**DOI:** 10.11889/j.1000-3436.2021.rj.39.050101

## 溶剂化电子的传奇历史

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**摘要** 在科学发展的过程中,存在着大量的基础性实验,它们的进一步发展构成了成熟的学科体系。比较罕见的是,来自不同领域的知识在某一时刻碰撞所产生的火花给以往的成果提供了一个全新的视角,并引发了新一波的研究浪潮。溶剂化电子就是这样的情况,它是一个在液体中孤立且电荷极化的电子,那么,这样一个概念是源自何处呢?

**关键词** 溶剂化电子,性质,结构模型,金属-液氨溶液,辐射化学

**中图分类号** TL13

### 1 Alkaline metal solutions in liquid ammonia

In the mid-19th century, liquid ammonia was found to be able to dissolve without reaction, unlike water, sodium and other alkaline metals in intense blue solutions that were widely used, notably by Weyl<sup>[1]</sup>, to effect previously difficult chemical reductions of organic or inorganic molecules<sup>[2]</sup>. However, the handling of this solvent is very complex because it is liquid only below  $-33^{\circ}\text{C}$  at normal pressure. It must therefore be kept cold, but also sheltered from the atmosphere, otherwise sodium would be oxidized to the air and cold ammonia would

condense water vapours. In addition, metal solutions are metastable and, even very pure, they decompose very slowly into hydrogen and amide. Despite this, Kraus, assistant to A. A. Noyes at the Massachusetts Institute of Technology, decided in 1904 for his thesis subject to measure the conductivity of various ions in salt solutions in liquid  $\text{NH}_3$  (Fig. 1)<sup>[3]</sup>.

While he had already determined the conductivity of each ion of alkaline salts, Kraus also wanted to compare these results with measurements in alkaline metal solutions and, to his surprise, their conductivity was far superior to those of salts. Stranger also, conductivity did not depend on the

收稿日期: 初稿 2021-06-28; 修回 2021-07-12

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Received 28 June 2021; accepted 12 July 2021

nature of the metal. He therefore concluded in 1908<sup>[4]</sup> first that alkaline metals were dissociated in ammonia into ions, and then that the common complementary anion of these solutions, responsible for most of their conductivity and blue color, could only be a solvated electron: "the anion is an electron surrounded by solvent molecules" ... "a solvated electron"...

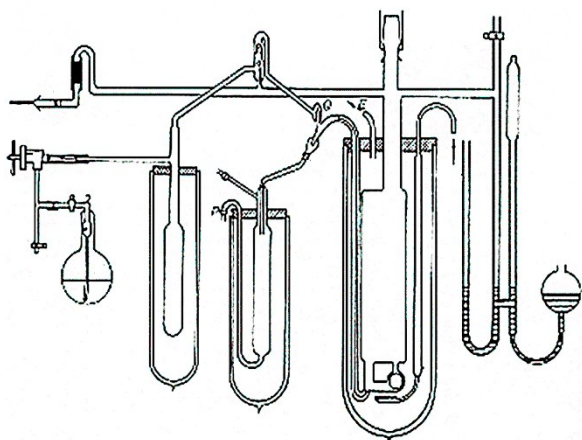


**Fig.1** C.A. Kraus vacuum equipment to measure the conductivity of alkaline metal solutions in low-temperature liquid ammonia<sup>[3]</sup>.

This audacious concept of a solvent electron, similar, despite its very different nature, to a monovalent anion stabilized by the polarization of solvent molecules, immediately prompted a great deal of work<sup>[2]</sup>. Gibson and Argo established optical absorption spectra for the solvated electron in diluted alkaline and alkaline-earthous metals solutions in ammonia and methylamine<sup>[5-6]</sup>. They were characterized by a wide, intense and asymmetrical band in energy, with a maximum in the near IR. Later, it was shown that at lower temperature or higher pressure these spectra were moved to the visible<sup>[2]</sup>, as were the halide absorption spectra. Like them, they have been attributed to a transfer of the charge to the molecules of the solvation layer (or CTTS, charge transfer to solvent). Early on, the expansion of the volume of the solution observed when the metal dissolved in the ammonia had suggested that the electron occupied a cavity much larger than its own size. This result inspired the early theoretical models of the structure of the solvated electron described by Ogg<sup>[7]</sup> and then by Jortner<sup>[8]</sup>. The spectrum was attributed to an excitation of the  $1s \rightarrow 2p$  transition of

a particle, in a cavity created by the mutual repulsion of hydrogen atoms of the polar solvent molecules oriented towards the electron.

## 2 Ionizing radiation

Röntgen's<sup>[9]</sup> discovery in December 1895 of an invisible and very penetrating X-rays, generated by the impact of cathode rays on the anticathode of a Crookes tube, and then by H. Becquerel of the uranic rays in 1896, suddenly extended the fields of photophysics and photochemistry to much more energetic radiation. Pierre and Marie Curie discovered in the pechblende in 1898 two new elements much more radioactive than uranium, polonium<sup>[10]</sup> and radium<sup>[11]</sup>, and emitting specific radiation (Fig. 2).

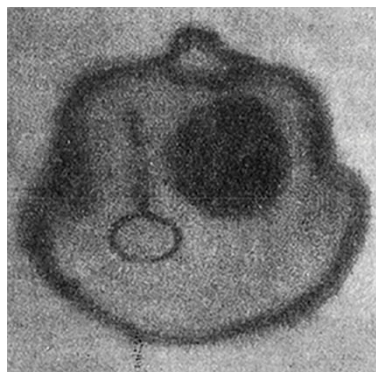


**Fig.2** Image of a wallet on a photographic plaque in Marie Curie's thesis<sup>[12]</sup>. The radiation used was  $\gamma$  rays emitted by a source of radium contained in glass (the radiation  $\beta$  was deflected by a magnetic field). The mode of penetration is very similar to that of X-rays.

The following year, they published the physical-chemical effects of these radiations on matter<sup>[13]</sup>, in particular, in addition to the ionization of gases, the production of ozone from oxygen, and the production of molecular hydrogen and hydrogen peroxide in radioactive water solutions. The initial ions of the water are assumed to be  $H_2O^+$  and  $H_2O^-$  which, after a fast reaction with  $H_2O$ , give birth respectively to the radicals  $OH\cdot$  and  $H\cdot$ <sup>[14]</sup>. On the basis of this radical model, the radiolytic yields of products formed in various solutions can be gradually explained<sup>[15-18]</sup>.

## 3 The hydrated electron

However, in the 1950s, some results were beyond this<sup>[19-20]</sup>. In particular in 1952, G. Stein of the

Hebrew University of Jerusalem found very different results for the reduction of aqueous methylene blue solutions in acidic or neutral environments<sup>[21]</sup>. He concludes that in addition to the radical  $H^\bullet$ , a second reductive species should be considered. He was the first to propose the hydrated electron  $e_{aq}^-$  as an analog of the electron ammonized  $e_{am}^-$ , known in solutions of alkaline metals in liquid ammonia. The hydrated electron would come from the electron originally derived from the ionization of the water. Immediately, the theorist R. L. Platzman<sup>[21]</sup> describes in his model how this electron can escape recombination with the parent cation and thus solve the solvent.

Despite this, the hypothesis of a hydrated electron will take a decade to impose itself. The first argument against it was the stability of  $e_{am}^-$  while  $e_{aq}^-$  would be a very short-lived intermediary. Nor could any metastable blue color of the ammoniated electron be observed after liquid ammonia irradiation by a continuous accelerator<sup>[22]</sup>. Prudently, the authors concluded that his concentration may have been too low. Another difficulty was that the aqueous solutions studied in radiolysis, even deaerated, were often very acidic and in this case the hydrated electron, reacting with a proton, is actually replaced by a radical  $H^\bullet$ . In addition, even in a neutral environment, many solutes are reduced indiscriminately by  $H^\bullet$  or  $e_{aq}^-$  by giving the same products.

On the other hand, G. Stein and J. Jortner observe a fleeting blue color by putting water or alcohol in contact with sodium under argon atmosphere, suggesting the existence, at least transient, of a solvated electron also in these liquids<sup>[23]</sup>. In 1958, chloroacetic acid radiolysis produced hydrogen in an acidic environment but chloride ions in a neutral environment<sup>[24]</sup>. In 1961, studies of the ionic force on the radiolysis of cationic or anionic solutes show that the predominant radical species in neutral environment carries a negative charge<sup>[25]</sup>.

However, in a 1961 review of water radiolysis by Hart and Platzman, the hydrated electron is not even mentioned. The resolution of this controversy is

also not the objective of the early experiments to detect reactional intermediaries by the pulsed radiolysis method developed by Matheson and Dorfman who observe the free radical  $I_2$ <sup>[26]</sup>.

Similarly, J. W. Boag<sup>[27]</sup>, in his relationship of historical observations of  $e_{aq}^-$  using the pulsed radiolysis installation he had just developed at the Gray Laboratory in London, reports that E. J. Hart had in fact come to join him in order to detect free carbonate radicals. But two distinct absorption maxima are observed, one of which is found in pure water, and is moved to the infrared when the water is added with ammonia. The spectrum is therefore assigned to  $e_{aq}^-$ . The key result of this highlight was communicated to the 2nd International Congress of Radiation Research in Harrogate in 1962. At the same conference, J. P. Keene also reports on his observation, independently, of the new spectrum by using the pulse radiolysis installation he built at the Patterson Laboratory in Manchester, spectrum which he also assigned to the hydrated electron. Boag proposes to publish their respective results simultaneously in the same issue of the journal *Nature*<sup>[28-30]</sup>. In the same year, J. Jortner, Ottolenghi and G. Stein observe by flash photolysis the same spectrum of  $e_{aq}^-$  produced by photodetachment from the iodide anion<sup>[31]</sup>.

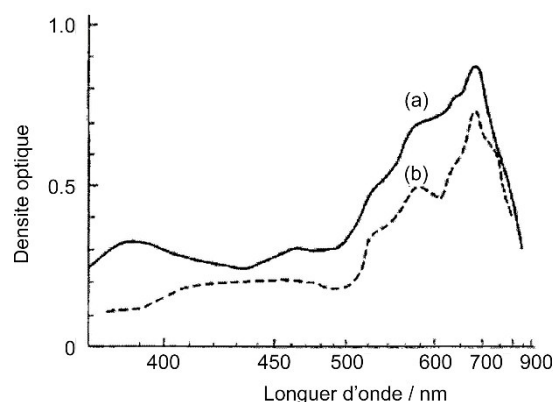


**Fig.3** Optical absorption spectrum of hydrated electron observed for the first time by pulse radiolysis in a solution of sodium carbonate (with the peaks of the radical  $CO_3^{\bullet -}$  at 600 nm and  $e_{aq}^-$  at 700 nm) (a), and in pure water (with  $e_{aq}^-$ ) (b), both deaerated<sup>[31]</sup>. To detect a nanomolar concentration of  $e_{aq}^-$ , the optical path was elongated by placing the 4 cm optical cell at the center of a multiple reflection system. At this concentration, the decay of  $e_{aq}^-$  by recombinations is very slow and is observed for more than 50 ms.

#### 4 The solvated electrons

These results had a great impact and triggered a proliferation of research. A Weyl Symposium was created, with its first edition in 1963<sup>[32]</sup>, to commemorate the centenary since Weyl's discovery of metal solutions in ammonia. In addition to the recent works by specialists in the field, Jortner<sup>[32]</sup> presents a review on the evolution of theories on the structure of the solvated electron that preceded its semi-continuous model, which he applied not only to  $e_{am}^-$  but also for the first time to  $e_{aq}^-$  which had just been highlighted. This structure consisted of a cavity surrounded by a few polarized solvent molecules and then a continuum of solvent, characterized in particular by its density and dielectric constants, optical and static. The model accounts for the wavelengths of the respective maxima according to temperature and pressure. Very quickly, thanks to newly installed pulse radiolysis facilities and shorter pulses, a very wide variety of liquids, more or less polar, were ionized and solvated electrons detected. Despite their instability, the very similar properties of their transient absorption spectra were compared with fast IR detections in alcohols and polyols, amines, ethers, sulphides, carbonates, ionic liquids and even hydrocarbons<sup>[33-36]</sup> (Fig. 4), or their reactivity with series of molecules.

Hereafter, the following editions of the Weyl Symposiums<sup>[36]</sup> have brought together experimentalists and theorists discussing solvated electrons produced either by pulsed radiolysis or by dissolution of alkaline and alkaline-earth metals in liquid ammonia, amines and certain ethers. One of the interesting aspects of the solvated electron is to appear thus as the smallest solvated chemical species which, despite its transient nature, can be formed and studied in the greatest number of environments. It is also compared with electrons trapped in irradiated ices<sup>[37]</sup>.

Thanks to the developments of the simulation, numerous theoretical models have been proposed to account for the properties of the solvated electron, including the hydrated electron, and to describe its

Structure (Fig.5)<sup>[38]</sup>.

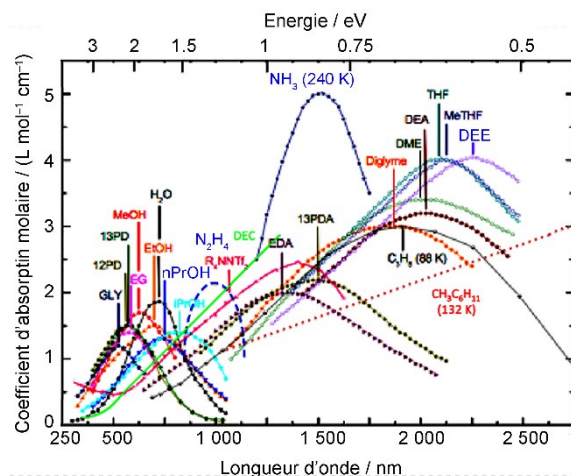


**Fig.4** Optical absorption spectra of electrons solvated in different liquids, calibrated into molar absorption coefficients. GLY: glycol; PD: propane diol; EG: ethylene glycol; MeOH: methanol; EtOH: ethanol; PrOH: propanol; EDA: ethane-1,2-diamine; DEA: diethanolamine; 13PDA: propane-1,3-diamine; THF: tetrahydrofurane; DME: dimethylether; DEE: diethylether; Diglyme: bis(2-methoxyethyl)ether; R4NNTf2: methyl-tributyl-ammonium bis(trifluoromethyl-sulfonyl)imide; DEC: diethylcarbonate. (Adapted from<sup>[33-36]</sup>).

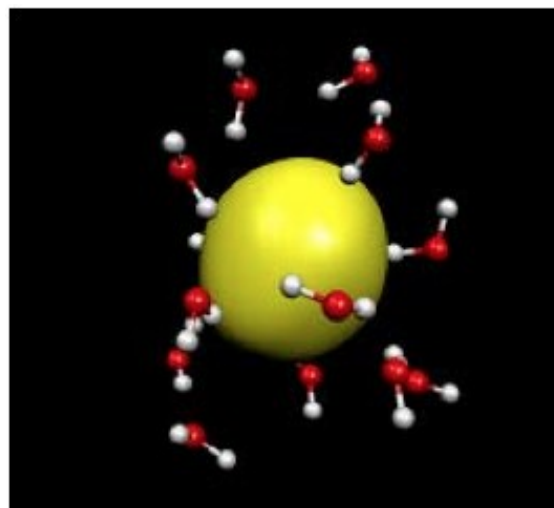


**Fig.5** Structure of the hydrated electron simulated by molecular dynamics<sup>[38]</sup>. The charge of the electron is delocalized in the 2.5 Å radius cavity. H<sub>2</sub>O molecules are oriented towards the central charge by one of their H atoms.

The perturbations of water molecules in the solvation layer are sometimes calculated<sup>[39]</sup>. A recent review analyzed these models in detail. The simulated optical absorption spectrum that is closest to the experiment is composed of transitions between the fundamental state and about fifteen excited states (Fig.6)<sup>[40]</sup>.

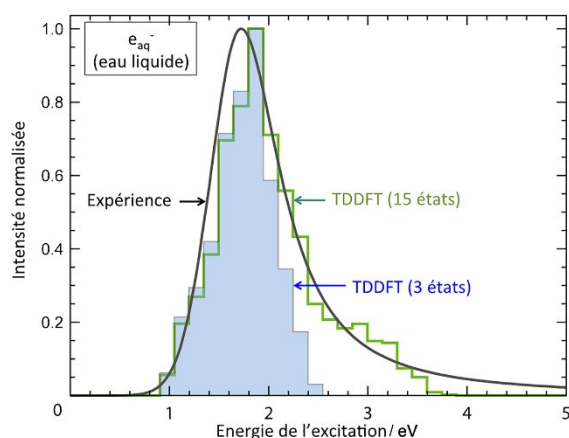


**Fig.6** Comparison between the optical absorption spectra of the hydrated electron obtained by the experiment or based on simulations based on the theory of the functional of density with 3 or 15 states<sup>[40]</sup>.

## 5 Box 1 - Instability of $e_{\text{solv}}^-$

How to explain the large difference between the  $e_{\text{am}}$  metastability in blue metal solutions in ammonia or amines, and the rapid decay of solvated electrons produced by pulse radiolysis, including in  $\text{NH}_3$  (which explains why the 1953 experiments failed to observe any blue color<sup>[24]</sup>)? In fact, the solvated electron is in a fundamental state, which would be stable if its existence were not limited by its very high chemical reactivity that depends on the conditions (and improperly characterized by a lifetime). In  $\text{NH}_3$  and amines, it is inert against alkaline or alkaline-earth cations. It therefore decreases only very slowly with impurities or acidic ions slowly released from the walls of the container (the ionization product of  $\text{NH}_3$  itself is only  $K_{\text{ion}} = 10^{-23} \text{ M}^2$ ). While in radiolysis,  $e_{\text{am}}^-$  reacts with the  $\text{NH}_4^+$  cations and other radical species formed simultaneously. In addition, it was shown by pulse radiolysis that in water the dismutation reaction between two  $e_{\text{aq}}^-$  was very fast in producing dihydrogen and  $\text{OH}^-$ . It also occurs during the contact between water and sodium metal and produces hydrogen explosions with oxygen from the air, well known to apprentice chemists.

But, provided that liquid ammonia is added with hydrogen to capture oxidizing radicals, and Alkaline amidurn  $\text{NaNH}_2$  to capture  $\text{NH}_4^+$  ions, the radiation-induced electron  $e_{\text{am}}$  becomes as stable as in metal solutions. Thanks to this inertia, the characteristic

blue color was observed in this case after a simple stationary irradiation (Fig. 7)<sup>[41]</sup>, and also after photolysis and photodetachment from  $\text{NH}_2^-$ . Whatever the mode of formation, the identity between the solvated electrons was thus definitively established. Blue  $e_{\text{am}}$  solutions were also obtained in  $\text{NH}_3$  by electrochemistry (with alkaline halide as the electrolyte)<sup>[2, 42]</sup> or by photo-injection from a semiconductor<sup>[43]</sup>.

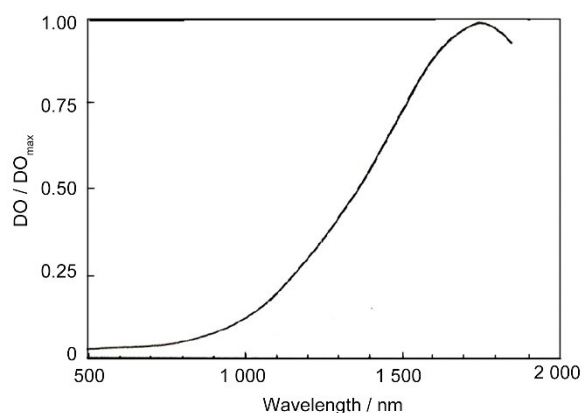


**Fig.7** Optical absorption spectrum of a hydrogen and sodium amide solution in liquid ammonia irradiated at 20 °C (spectrum is recorded in a sealed test tube, resistant to 91.19 kPa<sup>[41]</sup>).

A.J. Swallow<sup>[44]</sup> even calculated that, given the formation of  $e_{\text{aq}}^-$  by irradiation by cosmic rays and by photodetachment from chlorides by the Sun's UV rays, its near-stationary concentration in the oceans would be about  $10^{-9}$  molar ("Many of the short-lived chemical species ..... are known to exist naturally.....") (Fig.8).

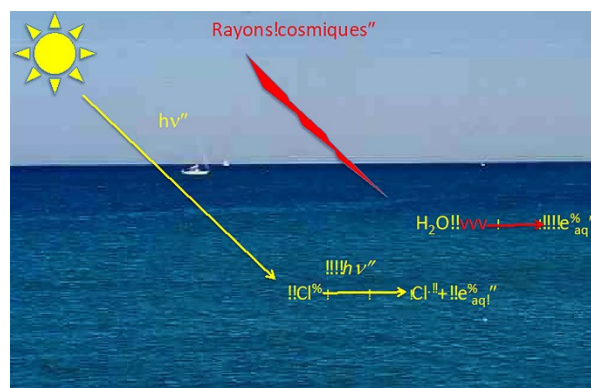


**Fig.8** Formation of  $e_{\text{aq}}^-$  in the oceans<sup>[44]</sup>.

As early as 1971, Baxendale and Wardman<sup>[45]</sup> observed for the first time after a 5 ns pulse the solvation dynamics of an electron in the viscous *n*-propanol at low temperature, which slowed down

molecular movements (Fig.9). The initial spectrum is located in the infrared and the absorption at 1 300 nm then disappears in 1  $\mu$ s while simultaneously the spectrum of the completely solvated electron develops and stabilizes with its maximum at 500 nm. The first spectrum is assigned to a pre-solvated electron and the solvation time is estimated at  $\tau_{1/2}=50$  ns (Fig.9, insets).

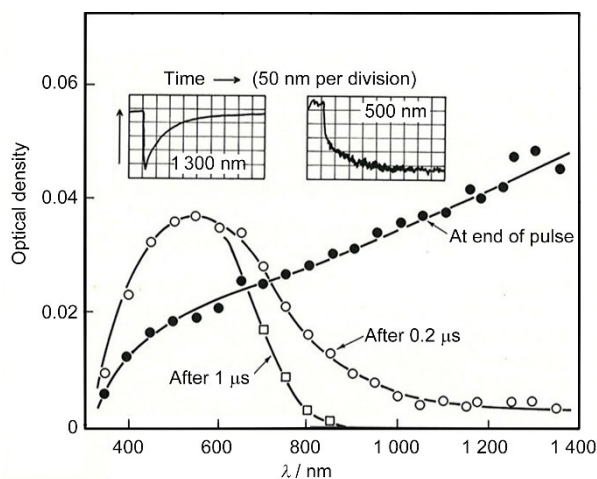


**Fig.9** Solvation of  $e_s^-$  in the n-propanol at 152 K after ionization by a 5 ns pulse of a pulse accelerator<sup>[45]</sup>. Insets: pre-solvated electron decay at 1 300 nm and growth of  $e_s^-$  at 500 nm.

It will be necessary to wait for the pulses of 100 femtoseconds offered by the lasers with a biphotonic UV excitation for the electron solvation, much faster at room temperature, can be observed in water ( $\tau_{1/2}=200$  fs)<sup>[46]</sup> (Fig. 10) or polyols ( $\tau_{1/2}=10$  ps for glycerol)<sup>[47]</sup>. In liquid ammonia at  $-50$  °C, solvated electrons in blue solutions were excited in the IR band at 1 280 nm to eject them from their solvent cavity and the relaxation time to the solvent state was  $\tau_{1/2}=150$  fs<sup>[48]</sup>, which, given the temperature, is even faster than in the water. On the other hand, the solvation relaxation of an electron in the ionic liquid  $[C_4\text{mpyr}][\text{NTf}_2]$ , with larger molecules and heterogeneous density domains, is distinguished by two different velocity processes ( $\tau_{1/2}=70.4$  ps and  $\tau_{1/2}=574$  ps)<sup>[49]</sup>.

Gradually, advances in pulse radiolysis of water solutions have helped to determine the UV optical absorption spectra of  $H^-$  and  $OH^-$  radicals. But after subtracting these bands from the total spectrum in pure water, it remained another component, correlated

with the infrared absorption band of  $e_{aq}^-$ , and shifted like it between 20 and 34 °C to the larger wavelengths, but much less intense (Fig. 11, inset)<sup>[50]</sup>. With this precise deconvolution, the authors concluded that the UV component must be assigned to  $e_{aq}^-$  and that it corresponds to the edge of the absorption band, shifted to the visible, of water molecules from the solvation layer, disturbed by the transfer of the electron charge. This would therefore be the first observation of a spectrum of a solvation layer.

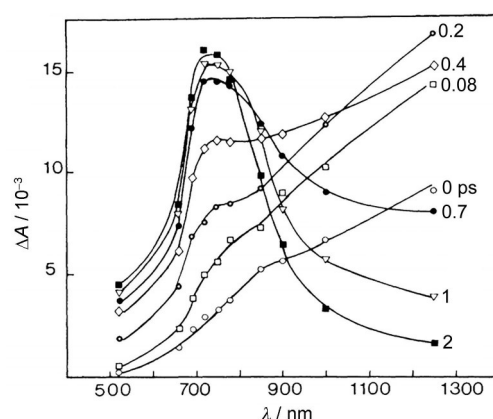


**Fig.10** Solvation of  $e_{aq}^-$  after ionization of water molecules at 294 K by a biphotonic UV excitation by a laser pulse of 100 fs ( $\tau_{1/2}=200$  fs)<sup>[46]</sup>.

Given the very comparable properties of all solvated electrons, the UV absorption spectra of very pure sodium solutions in ammonia were in turn examined at  $-50$  °C and 20 °C<sup>[51]</sup>. In this case, the solution contains only  $e_{am}^-$  and  $Na^+$ . The spectra display indeed UV absorption bands proportional to the concentration of the metal and also correlated with the IR bands of  $e_{am}^-$  (Fig.11).

They are shifted from the band edge of pure ammonia, as well as those of  $e_{aq}^-$  at 20 °C and 34 °C compared to the pure water band edges (Fig. 11, inset). These elements therefore support the previous interpretation of a spectrum due to solvation molecules, which, by analogy, should also appear in the spectra of all other solvated electrons. More generally, the 'solvation' absorption band should exist for ions whose spectrum is also due to a charge transfer to solvent, such as halides, but in this case it would be masked by the main UV spectrum.

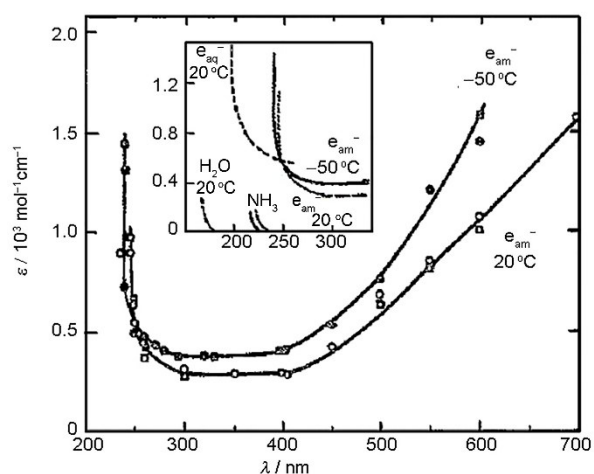


**Fig.11** UV-visible optical absorption spectra of the ammoniated electron at  $-50\text{ }^{\circ}\text{C}$  and  $20\text{ }^{\circ}\text{C}$ . Inset: comparison between the UV optical absorption bands of  $e_{\text{am}}^-$  at  $-50\text{ }^{\circ}\text{C}$  and  $20\text{ }^{\circ}\text{C}$  and  $e_{\text{aq}}^-$  at  $20\text{ }^{\circ}\text{C}$  and the  $\text{NH}_3$  and  $\text{H}_2\text{O}$  solvents; scales multiplied by 4; adapted from [50] and [51].

## 6 Conclusion

A consensus seems to emerge that the solvated electron occupies a cavity of size and shape varying with temperature and pressure, that it attracts towards its negative charge the H atoms of one of the polar groups of the solvation molecules whose structure is perturbed, and that its optical spectrum, very dissymmetrical in energy, results from a transition between a fundamental state and several linked excited states. Actually, the solvated electron remains a particularly interesting probe of solvation in all kinds of environments, and therefore of their structure, and this strange chemical species will probably spark much more research.

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#### Additional information of the author

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**Ph.D.** Doctorat d'Etat (Physical Sciences) (1961) University of Paris, at Radium Institute, (Head: Prof. Moïse Haissinsky). Ion adsorption on various surfaces by means of radioindicators.

#### Activity

1957–1970: Researcher at CNRS Radium Institute, Paris.

1970–1978: Researcher CNRS (senior in 1975) at Laboratoire de Physico-Chimie des Rayonnements (Head: Prof. Michel MAGAT), University Paris-Sud, Orsay.

1973 – 1988: Lectures for Master students in Molecular Physical Chemistry, University Paris-Sud, on Interaction of Ionizing Radiation with Matter.

1974–2002: Senior Researcher (CNRS) and leader of the group *Radiation Chemistry*.

1989–1998: Head of Laboratoire de Physico-Chimie des Rayonnements (75 persons). Until 2000: Head of Project ELYSE, Fast kinetics Center, inaugurated at Orsay in 2001 (Picosecond electron accelerator coupled with a femtosecond laser).

2002: Member as Senior Researcher Emeritus of Laboratoire de Chimie Physique created in 2000 by fusion of 3 laboratories including the Laboratoire de Physico-Chimie des Rayonnements.

#### Domains of research interest

##### Physical chemistry, Radiation Chemistry

Chemical mechanisms studied by pulse radiolysis and ultrafast time-resolved detection of transient species involved in irradiated solutions. Particularly, reactivity and properties of solvated electrons (optical absorption spectra, time-dependent yields, rate constants...) in various non aqueous solvents (ammonia, amines, ethers, alcohols, sulfides, carbonates...). Photoinjection by laser pulses of electrons solvated at semiconductor-electrolyte interface.

Radiolytic formation of long-lived ammoniated electrons at room temperature. *Nature*, 1971, **232**: 173.

Ultra-violet absorption (of solvated electrons) by metal-ammonia solutions. *Nature*, 1976, **263**: 47.

**Nucleation, growth mechanisms and nuclearity-dependent properties of clusters of metals and semiconductors** (absorption spectra, rate constants, nuclearity-dependent redox potential of metal nanoparticles different from that of bulk metal and demonstrated for the first time in 1973). Some of these nanoparticles have been studied in various industrial applications, such as catalysis, non linear optics, electronics and photography. Explanation of the auto-catalytic mechanism of the photographic development. Enhancement by a factor of  $\times 10$  of the photographic sensitivity by doping the silver halide crystals.

Micro-aggregates of non-noble metals and bimetallic alloys prepared by radiation-induced reduction. *Nature*, 1985, **217**: 344.

Enhanced yield of photoinduced electrons in doped silver halide crystals. *Nature*, 1999, **402**: 866.

Magic of the Image : Video CNRS 1989, 26 min.

**Homogeneous crystal nucleation-growth dynamics** induced by single laser pulse in supersaturated solutions (accretion rate constant, concentration of nuclei...).

Homogeneous nucleation-growth dynamics induced by single laser pulse in supersaturated solutions. *Crystal Growth & Design*, 2012, **12**: 5980.

Anisotropic Time-Resolved Dynamics of Crystal Growth Induced by a Single Laser Pulse Nucleation. *Crystal Growth & Design*, 2021, **21**: 799-909.

### International Expertise

Twice Chairman of an international Conference (1991: the Miller Conference on Radiation Chemistry, and 2004: the Gordon Research Conference on Radiation Chemistry).

Member of the Editorial Board of international and national Journals.

Expert for national and foreign projects (1984: One month mission for the World Bank at USTC Hefei, China. 1998 and 2006:

Expert for Department of Energy, USA. 2007: Expert for Japanese Department of Nuclear Energy. 2012: Expert for the Research Assessment Exercise KTH, Stockholm, Sweden).

### Awards

Silver Medal of CNRS (1983).

Invited Fellow by the Japanese Society for Promotion of Science (1983).

Two Grands prix by the French Académie des Sciences (1991 and 2000).

Bruylants Chair and Medal of the Louvain University, Belgium (1993).

Chevalier de l'Ordre National du Mérite, France (1995).

Marie Curie Medal of the Polish Society for Radioactive Research (1998).

Honorary Professor of University of Sciences and Technologies of Academy of Sciences of China, Hefei (1996).

Prix Robert Doisneau of the Photography Museum, Bièvres, France (2001).

Professor in Residence in the Erudite Program of Kerala Government, India (2010).

Distinguished Senior Fellow of the Société Chimique de France (2017).

### Bibliometry

Author or co-author of around 220 primary scientific publications, review articles in international journals, book chapters and patents.

Co-editor of two books and of several journal special issues. Direct supervision of 22 Ph.D. theses, 20 industrial contracts and 5 international programs.

More than 200 invitations in international Conferences, in Institutes for seminars, and by AIEA or CNRS for Training schools.

Dr Belloni was invited for one month in 1984 by the World Bank as an Expert for the pulse radiolysis project and a general report at the University of Science and Technology of China, Hefei. In particular, she had the privilege to have frequent and friendly discussions with Pr Yang Tsen Tong, Honorary Vice-President of the USTC, who defended his thesis in France, also at the Radium Institute with Pr M. Haïssinsky in 1949. After this mission, started long and fruitful collaborations, long-term exchange programs including theses, and frequent mutual visits between the colleagues of USTC and University Paris-Sud. A partnership agreement was signed in 1998. Dr Belloni was honored by the USTC in 1996 as a Honorary Professor of the University.

(相关阅读:J. Belloni 教授与中国科学技术大学辐射化学领域的合作经历,见封二封三)