It was initially believed that the noble gases could not form compounds due to their full valence shell of electrons that rendered them very chemically stable and unreactive. All noble gases have full s and p outer electron shells (except helium, which has no p sublevel), and so do not form chemical compounds easily. Because of their high ionization energy and almost zero electron affinity, they were not expected to be reactive.

The heavier noble gases have more electron shells than the lighter ones. Hence, the outermost electrons experience a shielding effect from the inner electrons that makes them more easily ionized, since they are less strongly attracted to the positively charged nucleus. This results in an ionization energy low enough to form stable compounds with the most electronegative elements, fluorine and oxygen, and even with less electronegative elements such as nitrogen and carbon under certain circumstances. These compounds are listed in order of decreasing order of the atomic weight of the noble gas, which generally reflects the priority of their discovery, and the breadth of available information for these compounds.

**Radon compounds**

Radon is not chemically inert, but its short half-life and the high energy of its radioactivity make it difficult to investigate its sole described fluoride, RnF₂, and its reaction products.

**Xenon compounds**

The first published report, in June 1962, of a noble gas compound was by Neil Bartlett, who noticed that the highly oxidizing compound platinum hexafluoride ionised O₂ to O²⁺. As the ionisation energy of O₂ to O²⁺ (1165 kJ mol⁻¹) is nearly equal to the ionisation energy of Xe to Xe⁺ (1170 kJ mol⁻¹), he tried the reaction of Xe with PtF₆. This yielded a crystalline product, xenon hexafluoroplatinate, whose formula was proposed to be \(\text{Xe}^+\text{[PtF}_6]^−\). It was later shown that the compound is actually more complex, containing both XeFPtF₅ and XeFPt₂F₁₁. This was the first real compound of any noble gas.

In September 1962, Howard Claasen reported the synthesis of a simple (two-element) noble gas compound, xenon tetrafluoride, by subjecting xenon and fluorine to a high temperature. In November 1962, Rudolf Hoppe of Universität Münster reported that xenon and fluorine can react to form xenon difluoride.

In recent years, further studies of XeF₄ have appeared, and several additional compounds of noble gases, particularly xenon, have been prepared, among them the xenon fluorides (XeF₂, XeF₆), oxyfluorides (XeOF₂, XeOF₄, XeO₂F₂, XeO₃F₂, XeO₂F₄) and oxides (XeO₂, XeO₃ and XeO₄). Xenon difluoride has been produced by the exposure of Xe and F₂ gases to sunlight, a result which eluded observation for 50 years. Xenon fluorides react with several other fluorides to form fluoroxenates, such as sodium octafluoroxenate (Na₂XeF₈²⁻), and fluoroxenonium salts, such as trifluoroxenonium hexafluoroantimoniate (XeF₃SbF₆⁻).

In terms of other halide reactivity, short-lived excimers of noble gas halides such as XeCl₂ are prepared in situ, and are used in the function of excimer lasers.

Recently, xenon has been shown to produce a wide variety of compounds of the type XeOₙX₂ where \(n\) is 1, 2 or 3 and
X is any electronegative group, such as CF₃, C(SO₂CF₃)₃, N(SO₂F)₂, N(SO₂CF₃)₂, OTeF₅, O(IO₂F₂), etc.; the range of compounds is impressive, similar to that seen with the neighbouring element iodine, running into the thousands and involving bonds between xenon and oxygen, nitrogen, carbon, boron and even gold, as well as perxenic acid, several halides, and complex ions.

A compound with an \((\text{Xe-Xe})\) bond has been reported, and is noteworthy: the compound \((\text{Xe}_2\text{Sb}_2\text{F}_{11})\) contains a \((\text{Xe–Xe})\) bond, and it is the longest element-element bond known (308.71 pm = 3.0871 Å).[11] Short-lived excimers of \((\text{Xe}_2\text{)}\) are reported to exist as a part of the function of excimer lasers.

**Krypton compounds**

Following the first successful synthesis of xenon compounds, synthesis of krypton difluoride \((\text{KrF}_2)\) was reported in 1963. Krypton gas reacts with fluorine gas under extreme forcing conditions, forming \((\text{KrF}_2)\) according to the following equation:

\[
\text{Kr} + \text{F}_2 \rightarrow \text{KrF}_2
\]

\((\text{KrF}_2)\) reacts with strong Lewis acids to form salts of the \((\text{KrF}^+)\) and \((\text{Kr}_2\text{F}_3^{3+})\) cations. The preparation of \((\text{KrF}_4)\) reported by Grosse in 1963, using the Claasen method, was subsequently shown to be a mistaken identification.

Krypton compounds with other than \((\text{Kr-F})\) bonds (compounds with atoms other than fluorine) have also been described. \((\text{KrF}_2)\) reacts with \((\text{B}(\text{OTeF}_5)_3)\) to produces the unstable compound, \(\text{Kr(OTeF}_5)_2\), with a krypton-oxygen bond. A krypton-nitrogen bond is found in the cation \([\text{HC≡N–Kr–F}]^+\), produced by the reaction of \((\text{KrF}_2)\) with \([\text{HC≡NH}]^+\ [\text{AsF}_6^-]\) below −50 °C.

**Argon compounds**

The discovery of \(\text{HArF}\) was announced in 2000. The compound can exist in low temperature argon matrices for experimental studies, and has also been studied computationally. Argon hydride ion \((\text{ArH}^+)\) was obtained in the 1970s and was identified in the Crab nebula, based on the frequency of its light emissions.

**Contributors**

- Wikipedia