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NOBLE GASES: A RESEARCH STUDY

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INTRODUCTION

The **noble gases** are a group of chemical elements with very similar properties: under standard conditions, they are all odorless, colorless, monatomic gases, with very low chemical reactivity. The six noble gases that occur naturally are helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and the radioactive radon (Rn).

For the first six periods of the periodic table, the noble gases are exactly the members of **group 18** of the periodic table. However, this no longer holds in the seventh period (due to relativistic effects): the next member of group 18, ununoctium, is probably not a noble gas. Instead, group 14-member ununquadium exhibits noble-gas-like properties.

The properties of the noble gases can be well explained by modern theories of atomic structure: their outer shell of valence electrons is considered to be "full", giving them little tendency to participate in chemical reactions, and it has only been possible to prepare a few hundred noble gas compounds. The melting and boiling points for each noble gas are close together, differing by less than 10 °C (18 °F); consequently, they are liquids over only a small temperature range.

Neon, argon, krypton, and xenon are obtained from air using the methods of liquefaction of gases and fractional distillation. Helium is typically separated from natural gas, and radon is usually isolated from the radioactive decay of dissolved radium compounds. Noble gases have several important applications in industries such as lighting, welding, and space exploration. A helium-oxygen breathing gas is often used by deep-sea divers at depths of seawater over 180 feet (55 m) to keep the diver from experiencing oxygen toxemia, the lethal effect of high-pressure oxygen, and nitrogen narcosis, the distracting narcotic effect of the nitrogen in air beyond this partial-pressure threshold. After the risks caused by the flammability of hydrogen became apparent, it was replaced with helium in blimps and balloons.

The lightest elements are hydrogen and helium, both theoretically created by Big Bang nucleosynthesis during the first 20 minutes of the universe in a ratio of around 3:1 by mass (approximately 12:1 by number of atoms).

Almost all other elements found in nature, including some further hydrogen and helium created since then, were made by various natural or (at times) artificial methods of nucleosynthesis, including occasionally by activities such as nuclear fission.

The three main isotopes of the element hydrogen are often written as H for protium, D for deuterium and T for tritium. This is in order to make it easier to use them in chemical equations, as it replaces the need to write out the mass number for each atom. It is written like this:

D_2O (heavy water)
Instead of writing it like this:
^2H_2O

HELIUM (HE)

With a name taken from the Greek *helios* for "sun", helium is the second lightest and second most abundant gas in the known universe (hydrogen being number one). Because of its scarcity in our atmosphere its existence was not suspected until spectroscopic measurements revealed an unknown element present in the sun. The discovery of helium is generally credited to Janssen and Lockyear in 1868.

As no helium compounds are known, this family of gases was once thought to be inert. In 1962 the first noble gas compound was prepared with xenon (see below). Still, helium only occurs in uncombined form and must either be extracted from the atmosphere by liquefaction of air or separated from deposits of natural gas. It is thought that some of the terrestrial helium is the product of the alpha decay of radioactive isotopes beneath the crust.

Helium is the only known element which cannot be converted to a solid simply by cooling. It has 98% the lifting power of hydrogen with none of the Hindenburg-type drawbacks and is still used today in airships.

Table 1.1:

Atomic Number:	2	**Atomic Radius:**	140 pm (Van der Waals)
Atomic Symbol:	He	**Melting Point:**	<-272.2 °C
Atomic Weight:	4.00260	**Boiling Point:**	-268.93 °C
Electron Configuration:	$1s^2$	**Oxidation States:**	--

Isotopes

Seven isotopes of helium are known: Liquid helium (He-4) exists in two forms: He-4I and He-4II, with a sharp transition point at 2.174K. He-4I (above this temperature) is a normal liquid, but He-4II (below it) is unlike any other known substance. It expands on cooling, its conductivity for heat is enormous, and neither its heat conduction nor viscosity obeys normal rules.

Uses

- as an inert gas shield for arc welding.
- a protective gas in growing silicon and germanium crystals and producing titanium and zirconium.
- as a cooling medium for nuclear reactors, and
- as a gas for supersonic wind tunnels.

ARGON (AR)

Argon is third in abundance in the earth's atmosphere (about 1% by volume). Rayleigh and Ramsay isolated and identified the gas in 1894. Like the rest of the noble gases, it is colorless, tasteless, and odorless. The name for argon is taken from the Greek *argos* for "inactive".

For years argon has been used in ordinary incandescent light bulbs to replace the oxygen that would otherwise shorten the lifetime of the filament. It is used in some types of welding where active atmospheric gases would interfere with the process. Argon is also used in various types of "black lights" or UV lamps since excitation of the gas produces a significant amount of ultraviolet radiation. A few curious compounds have been made with argon but they are not very stable.

Table 1.2:

Atomic Number:	18	**Atomic Radius:**	188 pm (Van der Waals)
Atomic Symbol:	Ar	**Melting Point:**	-189.35 °C
Atomic Weight:	39.948	**Boiling Point:**	-185.85 °C
Electron Configuration:	$[Ne]3s^2 3p^6$	**Oxidation States:**	--

Isotopes

Naturally occurring argon is a mixture of three isotopes. Twelve other radioactive isotopes are known to exist.

Uses

It is used in electric light bulbs and in fluorescent tubes at a pressure of about 400 Pa. and in filling photo tubes, glow tubes, etc. Argon is also used as an inert gas shield for arc welding and cutting, as blanket for the production of titanium and other reactive elements, and as a protective atmosphere for growing silicon and germanium crystals.

KRYPTON (KR)

Named from the Greek *kryptos* or "hidden", krypton is neither green nor a solid material that can defeat Superman. Rather it is another noble gas discovered in 1898 by Ramsay and Travers. It ranks sixth in abundance in the atmosphere. Krypton gas is used in various kinds of lights, from small bright flashlight bulbs to special strobe lights for airport runways.

As with the other noble gases, krypton is isolated from the air by liquefaction.

One of the naturally occurring non-radioactive isotopes of krypton, Kr-86 (17.3%) is used as the basis for the current international definition of the *metre*. One metre is 1,650,762.73 wavelengths of the red-orange spectral line of krypton-86.

Table 1.3:

Atomic Number:	36	**Atomic Radius:**	202 pm (Van der Waals)
Atomic Symbol:	Kr	**Melting Point:**	-157.38 °C
Atomic Weight:	83.80	**Boiling Point:**	153.22 °C
Electron Configuration:	$[Ar]4s^2 3d^{10} 4p^6$	**Oxidation States:**	--

Isotopes

Naturally occurring krypton contains six stable isotopes. Seventeen other unstable isotopes are recognized. The spectral lines of krypton are easily produced, and some are very sharp. While krypton is generally thought of as a rare gas that normally does not combine with other elements to form compounds, it now appears that the existence of some krypton compounds can exist. Krypton difluoride has been prepared in gram quantities and can be made by several methods. A higher fluoride of krypton and a salt of an oxyacid of krypton also have been reported. Molecule-ions of ArKr+ and KrH+ have been identified and investigated, and evidence is provided for the formation of KrXe or KrXe+.

Krypton clathrates are prepared using hydroquinone and phenol. ^{85}Kr can be used for chemical analysis by imbedding the isotope in various solids. During this process, kryptonates are formed. Kryptonate activity is sensitive to chemical reactions at the solution surface. Estimates of the concentration of reactants are therefore made possible. Krypton is used in certain photographic flash lamps for high-speed photography.

XENON (XE)

Also discovered by Ramsay and Travers in 1898, xenon (from the Greek *xenos* for "strange") is the rarest of the stable noble gases in the air. It is still recovered by liquefaction techniques and is widely used in strobe lamps.

In 1962 the first noble gas compound was produced by Neil Bartlett, combining xenon, platinum and fluorine. It is now possible to produce xenon compounds in which the oxidation states range from +2 to +8(!).

Table 1.4:

Atomic Number:	54	**Atomic Radius:**	216 pm (Van der Waals)
Atomic Symbol:	Xe	**Melting Point:**	-111.79 °C
Atomic Weight:	131.30	**Boiling Point:**	-108.12 °C
Electron Configuration:	$[Kr]5s^2 4d^{10} 5p^6$	**Oxidation States:**	--

Isotopes

Natural xenon is composed of nine stable isotopes. In addition to these, 20 unstable isotopes have been characterized. Before 1962, it had generally been assumed that xenon and other noble gases were unable to form compounds. Evidence has been mounting in the past few years that xenon, as well as other members of zero valance elements, do form compounds. Among the "compounds" of xenon now reported are sodium perxenate, xenon deuterate, xenon hydrate, difluoride, tetrafluoride, and hexafluoride. Xenon trioxide, which is highly explosive, has been prepared. More than 80 xenon compounds have been made with xenon chemically bonded to fluorine and oxygen. Some xenon compounds are colored. Metallic xenon has been produced, using several hundred kilobars of pressure. Xenon in a vacuum tube produces a beautiful blue glow when excited by an electrical discharge.

Uses

The gas is used in making electron tubes, stoboscopic lamps, bactericidal lamps, and lamps used to excite ruby lasers that generate coherent light. Xenon is used in the nuclear energy field in bubble chambers, probes, and other applications where a high molecular weight is of value. The perxenates are used in analytical chemistry as oxidizing agents. ^{133}Xe and ^{135}Xe are produced by neutron irradiation in air cooled nuclear reactors. ^{133}Xe has useful applications as a radioisotope. The element is available in sealed glass containers of gas at standard pressure. Xenon is not toxic, but its compounds are highly toxic because of their strong oxidizing characteristics.

RADON (RN)

Discovered in 1900 by Friedrich Dorn, radon is a radioactive noble gas now regarded as a potential health hazard in some homes. It also has medical applications for cancer treatment. Its original name was to be *niton* for "shining" but it was eventually named as a derivative of radium. Radon is found in underground deposits where is it produced by uranium and radium decay.

Radon fluoride (RnF) has been produced and the compound glows with a yellow light in the solid state.

Table 1.5:

Atomic Number:	**86**	**Atomic Radius:**	**220 pm (Van der Waals)**
Atomic Symbol:	**Rn**	**Melting Point:**	-71 °C
Atomic Weight:	**222**	**Boiling Point:**	**-61.7 °C**
Electron Configuration:	**[Xe]6s2 4f145d106p6**	**Oxidation States:**	

Isotopes

Twenty isotopes are known. Radon-22, from radium, has a half-life of 3.823 days and is an alpha emitter; Radon-220, emanating naturally from thorium and called thoron, has a half-life of 55.6 s and is also an alpha emitter. Radon-219 emanates from actinium and is called actinon. It has a half-life of 3.96 s and is also an alpha emitter. It is estimated that every square mile of soil to a depth of 6 inches contains about 1 g of radium, which releases radon in tiny amounts into the atmosphere. Radon is present in some spring waters, such as those at Hot Springs, Arkansas.

Uses

Radon is still produced for therapeutic use by a few hospitals by pumping it from a radium source and sealing it in minute tubes, called seeds or needles, for application to patient. This practice has been largely discontinued as hospitals can get the seeds directly from suppliers, who make up the seeds with the desired activity for the day of use.

PHYSICAL PROPERTIES

1. State. All of them are mono atomic, colourless, odourless and tasteless gases.

2. Solubility. They are sparingly soluble in water. The solubility generally increases with increase in atomic number.

3. Boiling Point and Melting Point. Due to weak intermolecular van der Waal's forces between them they possess very low b. pt. and m. pt. in comparison to those of the substances of comparable atomic and molecular masses. However, the b. pt. and m. pt. increase with increase in atomic number because van der Waal's forces become stronger with increase in size of the atoms or molecules. Therefore, among noble gases radon has the highest m. pt. and b. pt. whereas He has the least m. pt. and b. pt.

4. Liquefaction. It is extremely difficult to liquefy these gases as there are only weak van der Waal's forces which hold atoms together. Since these forces increase with the increase in atomic size and population of electrons, ease *of liquefaction increases down the zero groups from He to Rn.*

5. Atomic Radii. In the case of noble gases, the atomic radii correspond to van der Waal' radii. Therefore, these are quite large as compared with atomic radii of the other atoms belonging to the same period. As we go down the group, the van der Waal's radius increases due to the addition of new electronic shells and increase in screening effect.

6. Ionisation Energies. The ionisation energies of noble gases are very high. This is due to the stable configurations of noble gases. However, the ionization energies decrease with increase in atomic number from He to Rn due to increase in atomic size and decrease in effective nuclear charge.

7. Electron Affinities. Due to the stable np^2np^6 electronic configurations, noble gas atoms have no tendency to accept additional electron. Therefore, their electron affinities are almost zero.

8. Enthalpy of Fusion and Enthalpy of Vaporization. In general, the enthalpies of fusion and the enthalpies of vaporization are low and increase down the group.

CHEMICAL PROPERTIES

The noble gases are generally inert and do not participate in the reactions easily. The inertness of noble gases is due to the following reasons:

(i) The atoms of noble gases have stable closed shell electronic configurations.

(ii) The noble gases have exceptiowi1y high ionization energies.

(iii) The noble gases have very low electron affinities.

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