

15.3: Development of the Modern Periodic Table

Learning Objectives

- Appreciate how the modern periodic table was constructed.
- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.
- Distinguish between the A (atomic mass) value and the Z (atomic number) value when viewing the periodic table.
- Compare and contrast Mendeleev's and Moseley's ordering of the periodic table.

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.

	4 werthig	3 werthig	2 werthig	1 werthig	1 werthig	2 werthig
	-	-	-	-	Li = 7,03	(Be =9,3?)
Differenz =	-	-	-	-	16,02	(14,7)
	C = 12,0	N = 14,04	O = 16,00	Fl = 19,0	Na = 23,05	Mg = 24,0
Differenz =	16,5	16,96	16,07	16,46	16,08	16,0
	Si = 28,5	P = 31,0	S = 32,07	Cl = 35,46	K = 39,13	Ca = 40,0
Differenz =	89,1 / 2 = 44,55	44,0	46,7	44,51	46,3	47,6
	-	As = 75,0	Se = 78,8	Br = 79,97	Rb = 85,4	Sr = 87,6
Differenz =	89,1 / 2 = 44,55	45,6	49,5	46,8	47,6	49,5
	Sn = 117,6	Sb = 120,6	Te = 128,3	J = 126,8	Cs = 133,0	Ba = 137,1
Differenz =	89,4 = 2*44,7	87,4 = 2*43,7	-	-	(71 = 2*35,5)	-
	Pb = 207,0	Bi = 208,0	-	-	(TI = 204?)	-

Figure 15.3.1: Copy of table on page 137 of (Julius) Lothar Meyer: "Die modernen Theorien der Chemie" (1864). (Copyright;By Kawarayaki - Own work, CC BY-SA 4.0, <https://commons.wikimedia.org/w/index.php?curid=75552705>)

Dimitri Mendeleev (1834–1907)

The father of the modern periodic table, Dmitri Mendeleev was born in Siberia. He was the youngest of fourteen living children. Shortly after Dmitri's birth, his father lost his sight. When Dmitri was thirteen, his father died suddenly. His mother was left to manage the family's glass factory to support the family. In 1848, their glass factory burned down leaving the family with little resources. Dmitri's mother chose to take her son to St. Petersburg, Russia. Here, he would study chemistry at his father's alma mater (Main Pedagogical Institute which is not called St. Petersburg University). After several illnesses and his mother's death, he would earn his doctorate in chemistry in 1865.

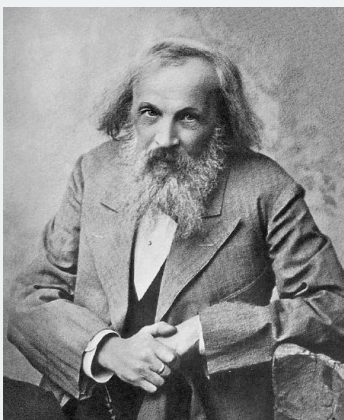


Figure 15.3.2 Dmitri Mendeleev, the father of the modern periodic table.

From 1867 to 1890, Dmitri served as a professor of general chemistry at St. Petersburg University. In his lifetime, he published more than 400 articles and books. Two of these works were noted in textbooks in general and organic chemistry.

As for his personal life, Dmitri Mendeleev married twice (his first marriage ended in a controversial divorce). Dmitri fathered a total of six children from both marriages. In 1907, he passed away from influenza complications. He was honored for his works by having element 101 (Mendelevium) named for him. Watch the video below to see how Mendeleev ordered his table of elements.



Mendeleev's Periodic Table

Mendeleev, who first published his periodic table in 1869 (Figure 15.3.1), is usually credited with the origin of the modern periodic table. The key difference between his arrangement of the elements and that of Meyer and others is that Mendeleev did not assume that all the elements had been discovered (actually, only about two-thirds of the naturally occurring elements were known at the time). Instead, he deliberately left blanks in his table at atomic masses 44, 68, 72, and 100, in the expectation that elements with those atomic masses would be discovered. Those blanks correspond to the elements we now know as scandium, gallium, germanium, and technetium.

Reihen	Gruppe I. — R ² O	Gruppe II. — RO	Gruppe III. — R ² O ³	Gruppe IV. RH ⁴ RO ²	Gruppe V. RH ³ R ² O ⁵	Gruppe VI. RH ² RO ³	Gruppe VII. RH R ² O ⁷	Gruppe VIII. — RO ⁴
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=40	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Figure 15.3.3: Mendeleev's Periodic Table, as Published in the German Journal Annalen der Chemie und Pharmacie in 1872. The column headings "Reihen" and "Gruppe" are German for "row" and "group." Formulas indicate the type of compounds formed by each group, with "R" standing for "any element" and superscripts used where we now use subscripts. Atomic masses are shown after equal signs and increase across each row from left to right.

The groups in Mendeleev's table are determined by how many oxygen or hydrogen atoms are needed to form compounds with each element. For example, in Group I (referred to as Gruppe 1), two atoms of each element (hydrogen, lithium, potassium) bond with

one atom of oxygen to form compounds. In Group VII (Gruppe VI), one atom of fluorine, F, chlorine, Cl, and bromine, Br, reacts with one atom of hydrogen to form respective compounds. According to Mendeleev, elements in the same column of the periodic table produced similar types of compounds. This trend could not predict the types of hydrogen/oxygen-based compounds resulting from the middle section of the periodic table. Mendeleev had difficulty explaining the chemical nature of this part of the table now referred to as the transition metals.

The most convincing evidence in support of Mendeleev's arrangement of the elements was the discovery of two previously unknown elements whose properties closely corresponded with his predictions (Table 15.3.1). Two of the blanks Mendeleev had left in his original table were below aluminum and silicon, awaiting the discovery of two as-yet-unknown elements, *eka-aluminum* and *eka-silicon* (from the Sanskrit *eka*, meaning "one," as in "one beyond aluminum"). The observed properties of gallium and germanium matched those of *eka-aluminum* and *eka-silicon* so well that once they were discovered, Mendeleev's periodic table rapidly gained acceptance.

When the chemical properties of an element suggested that it might have been assigned the wrong place in earlier tables, Mendeleev carefully reexamined its atomic mass. He discovered, for example, that the atomic masses previously reported for beryllium, indium, and uranium were incorrect. The atomic mass of indium had originally been reported as 75.6, based on an assumed stoichiometry of InO for its oxide. If this atomic mass were correct, then indium would have to be placed in the middle of the nonmetals, between arsenic (atomic mass 75) and selenium (atomic mass 78). Because elemental indium is a silvery-white *metal*, however, Mendeleev postulated that the stoichiometry of its oxide was really In_2O_3 rather than InO. This would mean that indium's atomic mass was actually 113, placing the element between two other metals, cadmium, and tin.

Table 15.3.1: Comparison of the Properties Predicted by Mendeleev in 1869 for *eka-Aluminum* and *eka-Silicon* with the Properties of Gallium (Discovered in 1875) and Germanium (Discovered in 1886)

Property	<i>eka-Aluminum</i> (predicted)	Gallium (observed)	<i>eka-Silicon</i> (predicted)	Germanium (observed)
atomic mass	68	69.723	72	72.64
element	metal	metal	dirty-gray metal	gray-white metal
	low mp*	mp = 29.8°C	high mp	mp = 938°C
	$d = 5.9 \text{ g/cm}^3$	$d = 5.91 \text{ g/cm}^3$	$d = 5.5 \text{ g/cm}^3$	$d = 5.323 \text{ g/cm}^3$
oxide	E_2O_3	Ga_2O_3	EO_2	GeO_2
	$d = 5.5 \text{ g/cm}^3$	$d = 6.0 \text{ g/cm}^3$	$d = 4.7 \text{ g/cm}^3$	$d = 4.25 \text{ g/cm}^3$
chloride	ECl_3	GaCl_3	ECl_4	GeCl_4
	volatile	mp = 78°C bp* = 201°C	bp < 100°C	bp = 87°C
*mp = melting point; bp = boiling point.				

One group of elements absent from Mendeleev's table is the noble gases, all of which were discovered more than 20 years later, between 1894 and 1898, by Sir William Ramsay (1852–1916; Nobel Prize in Chemistry 1904). Initially, Ramsay did not know where to place these elements in the periodic table. Argon, the first to be discovered, had an atomic mass of 40. This was greater than chlorine's and comparable to that of potassium, so Ramsay, using the same kind of reasoning as Mendeleev, decided to place the noble gases between the halogens and the alkali metals.

The Role of the Atomic Number in the Periodic Table

Despite its usefulness, Mendeleev's periodic table was based entirely on empirical observation supported by very little understanding. It was not until 1913, when a young British physicist, H. G. J. Moseley (1887–1915), while analyzing the frequencies of x-rays emitted by the elements, discovered that the underlying foundation of the order of the elements was by the *atomic number*, not the atomic mass. Moseley hypothesized that the placement of each element in his series corresponded to its atomic number Z , which is the number of positive charges (protons) in its nucleus. Argon, for example, although having an atomic mass greater than that of potassium (39.9 amu versus 39.1 amu, respectively), was placed *before* potassium in the periodic table. While analyzing the frequencies of the emitted x-rays, Moseley noticed that the atomic number of argon is 18, whereas that of

potassium is 19, which indicated that they were indeed placed correctly. Moseley also noticed three gaps in his table of x-ray frequencies, so he predicted the existence of three unknown elements: technetium ($Z = 43$), discovered in 1937; promethium ($Z = 61$), discovered in 1945; and rhenium ($Z = 75$), discovered in 1925.

📌 H. G. J. Moseley (1887–1915)

Moseley left his research work at the University of Oxford to join the British army as a telecommunications officer during World War I. He was killed during the Battle of Gallipoli in Turkey.



Figure 15.3.4: G.J. Moseley

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements). The word *halogen* comes from the Greek for “salt maker” because these elements combine with other elements to form a group of compounds called salts.

Each row of elements on the periodic table is called a **period**. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

The Periodic Table of the Elements

group 1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18

period 1 2 3 4 5 6 7

atomic mass
or most stable mass number
1st ionization energy
in kJ/mol

chemical symbol

name

electron configuration

atomic number

electronegativity

oxidation states
most common are bold

alkali metals
alkaline metals
other metals
transition metals
metalloids
nonmetals
halogens
noble gases
unknown elements
radioactive elements have masses in parenthesis

electron configuration blocks

notes

- as of yet, elements 113-118 have no official name designated by the IUPAC.
- 1 kJ/mol = 96.485 eV.
- all elements are implied to have an oxidation state of zero.

Figure 15.3.5: Types of Elements. Elements are either metals, nonmetals, or metalloids. Each group is located in a different part of the periodic table. An interactive Periodic table can be found here. (Public Domain; 2012rc).

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi-metal), as shown in Figure 15.3.5. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 15.3.5 metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids**. Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

✓ Example 15.3.1

Based on its position in the periodic table, classify each element below as metal, a nonmetal, or a metalloid.

- Se
- Mg
- Ge

Solution

- In Figure 15.3.5, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
- Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be metal.
- Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table display the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

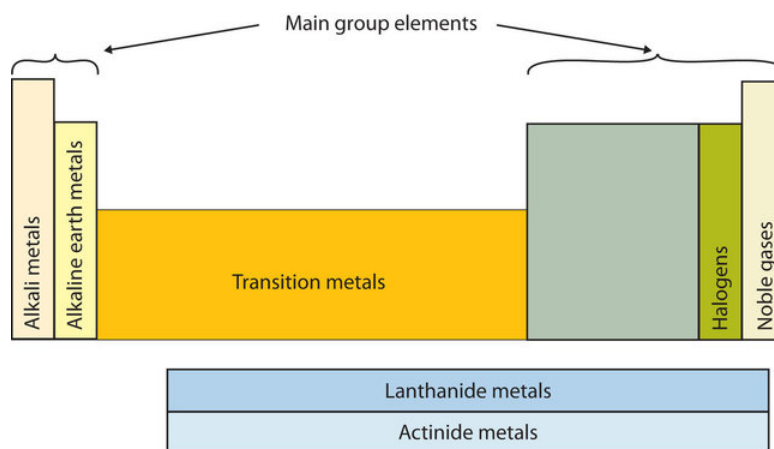


Figure 15.3.6: Some sections of the periodic table have specific names. Be sure to know the section names for quizzes and tests.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins. To watch a video of your instructor presenting these groups and periods, access your moodle class chapter 3 documents.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal. The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth's crust, respectively.

Group 2: The Alkaline Earth Metals

The alkaline earth metals are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 17: The Halogens

The halogens are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for "salt-forming," which reflects that all the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt). Compounds that contain fluorine ions (fluoride) are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth's ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for many years they were called inert gases or rare gases. However, the first chemical compounds containing noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

✓ Example 15.3.2: Groups

Provide the family/group names and period numbers (horizontal values) of each element.

- Li
- Ar
- Ra

Solution:

- Lithium is an alkali metal. It is located in period two.
- Argon is a noble gas. It is located in period three.
- Radium is an alkaline metal. It is located in period seven.

✓ Example 15.3.3: Classification of Elements

Provide elemental names for the following combinations:

- The alkali metal in period three.
- The halogen in period two
- A metalloid in period four
- A transition metal in period three

Solution:

- Sodium
- Fluorine
- Germanium or Arsenic
- There are no transition metals in period three (gotcha!)

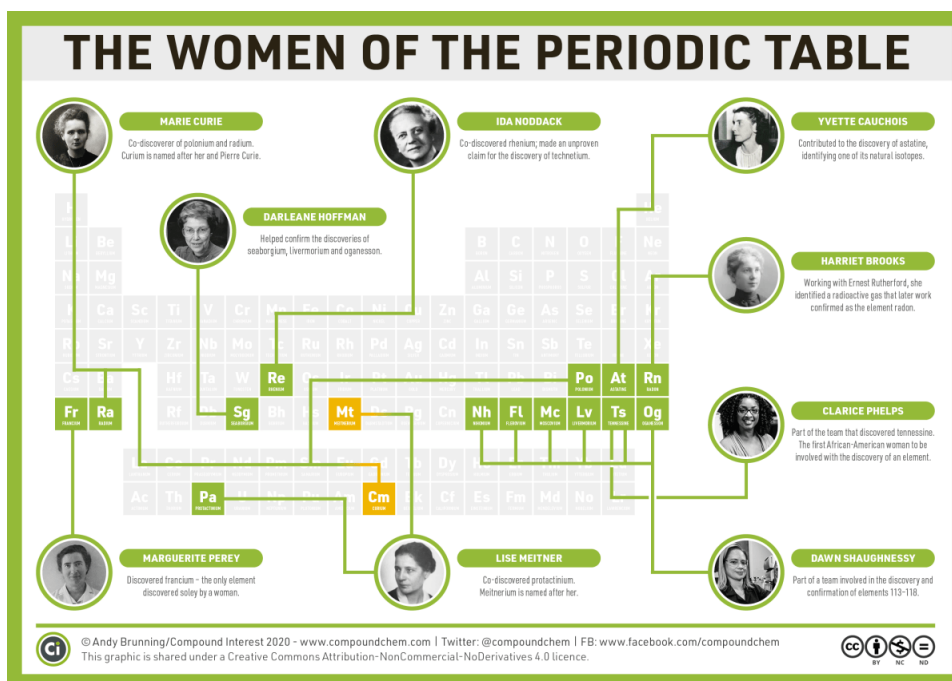


Figure 15.3.7: The Women of the Periodic Table Graphic by Andy Brunning.

References

- Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
- Sisler, Harry H. Electronic structure, properties, and the periodic law. New York: Reinhold publishing corporation, 1963.

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