Discovering Elements

CHEMISTRY • PERIODIC TABLE • DISCOVERING ELEMENTS

Section 1: The Formation of the Periodic Table

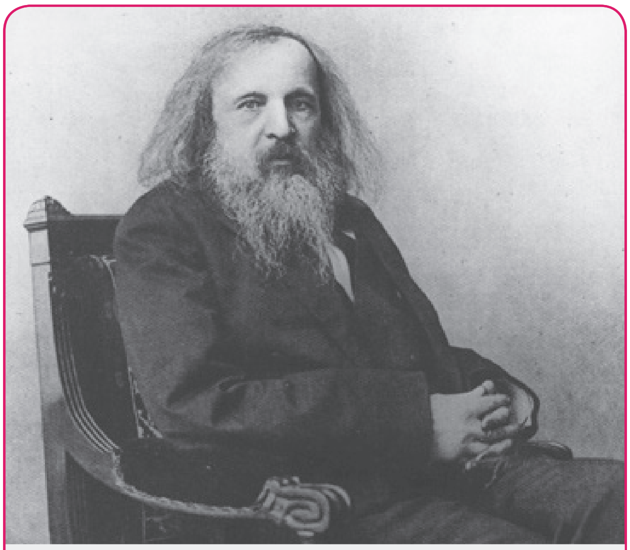
• Was Mendeleev the first person to notice patterns in the properties of the chemical elements?

The concept of a chemical element and an understanding of the nature of chemical change both took many centuries to develop, but by the end of the 18th century the role of oxygen in combustion had been established. Its predecessor, the phlogiston theory of combustion, had been decisively refuted, and Antoine Lavoisier (1743–1794) had produced the first modern list of the chemical elements.

From the late Middle Ages onwards, new elements had been discovered from time to time, sometimes during the investigation of metal ores, sometimes almost by accident, as in the case of Hennig Brand's discovery of phosphorus. In the 19th century more sophisticated techniques, such as spectroscopy and electrolysis, meant that more and more new elements were discovered and patterns in their properties began to be noticed.

For example, the alkali metals lithium, sodium and potassium, with relative atomic masses (then called atomic weights) of about 7, 23, and 39, respectively, share similar chemical properties. In 1829, Johann Döbereiner (1780–1849) pointed out that there also existed a numerical relationship between their atomic weights. 23, the relative atomic mass of sodium, is the mean of 7 and 39. He called this group of three elements a triad. Another triad appeared in the halogen elements chlorine, bromine and iodine, with relative atomic masses 35.5, 80 and 127, respectively. 80 is close to the mean of 35.5 and 127.

John Newlands (1837–1898) showed in 1863 that if the elements were arranged by relative atomic mass a pattern emerged, whereby elements seemed to share similar properties with those eight spaces along the table. Newlands termed this 'The Law of Octaves'. However, this musical analogy was met with scorn by other scientists, including one who suggested that such patterns were mere coincidences, just as you might get by placing the elements in alphabetical order. One critic pointed out that Newlands had assumed there were no elements left to discover, whilst another stated that similar elements, such as iron, cobalt and nickel, were too far apart in his table, yet dissimilar elements, such as platinum and bromine, were placed together. Although they had important insights, the ideas of Döbereiner and Newlands were thus not generally accepted at the time.



Dmitri Mendeleev published his first version of the periodic table in 1869

TABLE 01: Newlands' 1866 Periodic Table

H	F	Cl	Co/Ni	Br	Pd	I	Pt/Ir
Li	Na	K	Cu	Rb	Ag	Cs	Os
Be	Mg	Ca	Zn	Sr	Cd	Ba/V	Hg
B	Al	Cr	Y	Ce/La	U	Ta	Tl
C	Si	Ti	In	Zr	Sn	W	Pb
N	P	Mn	As	Di/Mo	Sb	Ni	Bi
O	S	Fe	Se	Rh/Ru	Te	Au	Th

Di is the symbol for the element "didymium", now known to be a mixture of the elements neodymium and praseodymium. Newlands also used the symbol "G" for beryllium, as it was then called "glucinum".

Extension Question

Q.1 Why were the ideas of Döbereiner and Newlands not accepted, but those of Mendeleev were?

The patterns noticed by Döbereiner and Newlands were regarded as mere coincidences, mainly because they could not explain why such patterns should exist. In the 1860s nothing was known about atomic structure – some famous scientists still doubted that atoms even existed – and so no theoretical reason for the patterns could be provided.

Mendeleev's achievement was the insight that there must be missing elements to fit in the gaps in his table, and the detail of his predictions. When gallium and germanium were discovered and found to have properties astonishingly close to his predictions, he became world-famous and his ideas became the basis for the modern classification of the elements.

• Suggested Films

- Discovery of Phosphorus
- The Curse of Phlogiston
- Phlogiston and Oxygen
- Introduction to the Periodic Table
- The Legacy of John Newlands

• What patterns in the elements did Mendeleev use in his periodic table?

Dmitri Mendeleev (1834–1927) published the first version of his periodic table in 1869. Like Newlands, he arranged the elements in order of their relative atomic masses, so that they fell into horizontal rows of similar elements called groups. In the modern periodic table these groups appear as columns. Sodium, potassium and lithium fell into one group, and chlorine, bromine and iodine into another. In order to preserve the essential feature of his table – elements with similar properties should be in the same group – he sometimes ignored the order of the relative atomic masses. In the case of tellurium and iodine, with relative atomic masses 128 and 127, respectively, he reversed the order in his table so that tellurium would fit into Group 6, beneath selenium, and iodine into Group 7, beneath bromine, as in both cases this is where their chemical properties fitted. We now know that the discrepancy arises because of the isotopes of tellurium and iodine.

He based his table on a vast range of chemical data: the properties of the elements, such as their densities, melting points and boiling points; the densities, melting and boiling points and formulae of their oxides and chlorides; and their acid-base properties. The numerical properties, such as melting points, showed a wave-like pattern rising and falling, with elements sharing similar properties appearing at regular intervals – hence the term 'periodic table'.

However, Mendeleev's greatest innovation was that, in order to keep similar elements in the same group, he deliberately left gaps for unknown elements, which he predicted would be discovered in the future, and about those properties he made precise predictions.

Extension Questions

Q2. Look at a modern periodic table. Can you find another example (other than tellurium and iodine) in which the order of relative atomic masses is NOT the same as the order of the atomic numbers. Why does this happen?

Potassium has relative atomic mass 39, but in the periodic table it comes after argon, with relative atomic mass 40. This happens because it is the atomic number (the number of protons in the nucleus), which decides the chemical nature of the atoms. Potassium has atomic number 19, and argon atomic number 18, making them fit into Group 1 and Group 0 respectively. The relative atomic masses of these elements are not in the same order as their atomic numbers because the weighted average mass of the isotopes of argon happens to be greater than the weighted average mass of the isotopes of potassium.

Q3. What causes the range of characteristics as you go down a column?

It is to do with the electron configuration of the atoms. As you go down a group, the full shells of electrons 'shield' the nucleus, and thus make it harder for the nucleus to attract electrons.

• Suggested Films

- Introduction to the Periodic Table
- What Is An Atom?

TABLE 02:
Timeline of the Development of the Periodic Table and Atomic Theory

The Development of the Periodic Table and Atomic Theory

CHEMISTRY • PERIODIC TABLE • DISCOVERING ELEMENTS

Dates	Development of chemical theory	Key people involved
c. 400 BCE	Earliest theories about atoms	Democritus (c. 460–c. 370 BCE)
c. 350 BCE	Greek four element theory	Empedocles (c. 490–c. 330 BCE)
1661	Early definition of 'element'	Robert Boyle (1627–1691)
1789	Modern definition of a chemical element; list of elements	Antoine Lavoisier (1743–1794)
1803	Connection made between atoms and chemical elements: each element contains only one type of atom	John Dalton (1766–1844)
1807	Isolation of potassium by electrolysis	Humphry Davy (1778–1844)
1813	Use of letters to stand for chemical elements e.g. 'H', 'O', etc.	Jöns Jacob Berzelius (1779–1848)
1829	Triads of elements	Johann Döbereiner (1780–1849)
1864	Patterns of 'octaves' in the elements	John Newlands (1837–1898)
1868	Discovery of helium in the Sun's spectrum	Jules Janssen (1824–1907), Norman Lockyer (1836–1920)
1869	Periodic table published, leaving gaps for undiscovered elements	Dmitri Mendeleev (1834–1927)
1871	Predictions made about unknown elements eka-aluminium, eka-silicon, eka-boron	Dmitri Mendeleev
1875	Discovery of gallium (= eka-aluminium)	Lecoq de Boisbaudran (1838–1912)
1879	Discovery of scandium (= eka-boron)	Lars Nilson (1840–1899)
1886	Discovery of germanium (= eka-silicon)	Clemens Winkler (1838–1912)
1894	Discovery of argon	Lord Rayleigh (1842–1919), William Ramsay (1852–1916)
1897	Discovery of the electron	J J Thomson (1856–1940)
1898	Discovery of neon, krypton, xenon	William Ramsay et al
1909–1911	Discovery of the atomic nucleus	Hans Geiger (1882–1945), Ernest Marsden (1889–1970), Ernest Rutherford (1871–1937)
1913	Theory of the hydrogen atom	Niels Bohr (1885–1962), Ernest Rutherford
1913	Discovery of isotopes	Frederick Soddy (1877–1956)
1919	Discovery of the proton	Ernest Rutherford
1932	Discovery of the neutron	James Chadwick (1891–1974)

TABLE 03:
Mendeleev's 1869 Periodic Table

			Ti	Zr	?
			V	Nb	Ta
			Cr	Mo	W
			Mn	Rh	Pt
			Fe	Ru	Ir
			Ni, Co	Pd	Os
H			Cu	Ag	Hg
	Be	Mg	Zn	Cd	
	B	Al	(eka-aluminium)	U	Au
	C	Si	(eka-silicon)	Sn	
	N	P	As	Sb	Bi
	O	S	Se	Te	
	F	Cl	Br	I	
Li	Na	K	Rb	Ca	Tl
		Ca	Sr	Ba	Pb
		(eka-boron)	Ce		
		Er	La		
		Yt	Di		
		In	Th		

The modern versions of the symbols of the elements have been used.
See Newlands' Table for the explination of 'Di'.
Eka-aluminium is now called gallium; eka-silicon is now called germanium;
eka-boron is now called scandium. Mendeleev did not use these names
or make exact predictions until 1871.

• What predictions did Mendeleev make about unknown elements?

In the fourth period of the modern periodic table, between calcium (Group 2) and arsenic (Group 5) there were no known elements that were similar to aluminium (Group 3) or to silicon (Group 4). Mendeleev, therefore, left gaps in his table in Group 3 and Group 4 for these missing elements. In 1871 he made precise numerical predictions about what these elements and their compounds. He called the element beneath aluminium, *eka-aluminium*, and the element beneath silicon, *eka-silicon*.

Properties	Mendeleev's predicitions about *eka*-aluminium	Properties of gallium
Relative atomic mass	68.8	69.72
Density (g/cm^3)	6.0	5.904
Melting point (°C)	Low	29.78
Formula of oxide	EA_2O_3	Ga_2O_3
Density of oxide (g/cm^3)	5.5	5.88
Nature of oxide	amphoteric - soluble in acids and alkalis	amphoteric - soluble in acids and alkalis
Formula of chloride	$EaCl_3$	$GaCl_3$

Gallium was discovered in 1875. Mendeleev's predictions about eka-aluminium were remarkably close to the actual values for gallium. Mendeleev became world-famous, and showed that the patterns in his periodic table were not accidents, they revealed deep truths about the nature of matter. It was not until the 20th century, however, that it became possible to explain these patterns in terms of the structure of the atom.

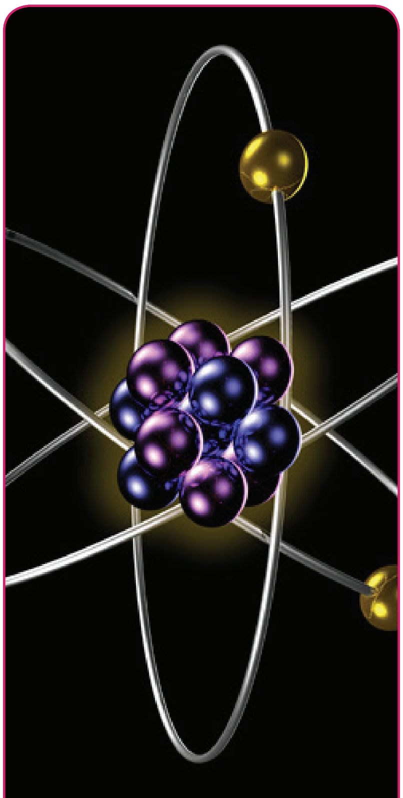
• Suggested Films

- **Atomic Structure**
- **Mendeleev's Prophecy**

• Suggested Activity

- **Below are Mendeleev's predictions for the element he called eka-silicon, a Group 4 element, which fitted into the gap he left below silicon. We now call this element germanium Ge. Research the properties of germanium and compare them with Mendeleev's predictions. How close are Mendeleev's predictions to the properties of germanium? What uses does the element germanium have? In what way are they similar to the uses of silicon?**

Properties	Mendeleev's predicitions about *eka*-silicon
Relative atomic mass	72
Density (g/cm^3)	5.5
Melting point (°C)	High
Formula of oxide	EsO_2
Density of oxide (g/cm^3)	4.7
Formula of chloride	$EsCl_4$
Boiling point of chloride (°C)	below 100



A 3D model showing the structure of the atom

Extension Question

Q4. Why was the discovery of gallium so important for Mendeleev's periodic table?

In the 19th century, no one knew for sure how many chemical elements existed. Mendeleev had taken a risk by leaving gaps in his table and by making such precise predictions. If there had turned out to be no elements to fit in the gaps, or if they had existed but their properties were completely different from those predicted by Mendeleev, his reputation would have been ruined.

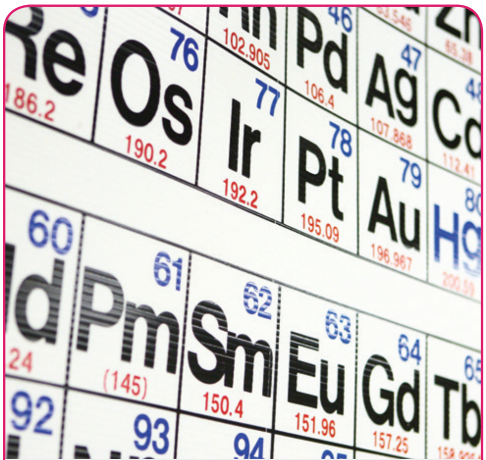
Section 2: Patterns in the Periodic Table

• How does the chemical nature of the elements change as we move from left to right across the periodic table?

If we look at Period 2 and Period 3, the elements change from being highly metallic in Group 1 (lithium, sodium) to become increasingly non-metallic through Groups 4, 5, 6, 7 and 0 (carbon, nitrogen, oxygen, fluorine, neon; silicon, phosphorus, sulphur, chlorine, argon). The oxides of the elements also change. Group 1 and Group 2 oxides (sodium oxide, calcium oxide) are strongly basic, and the oxides of Group 4 to Group 7 are mostly acidic when dissolved in water (carbon dioxide, sulphur dioxide, sulphur trioxide, nitrogen dioxide etc.).

• Suggested Films

- Atomic Structure
- Alkali Metals
- The Halogens
- The Noble Gases
- Reactivity Series



Part of the periodic table

Extension Question

Q5. Fluorine forms an oxide OF_2. Would you expect it to be acidic or alkaline in solution? Why?

The trend across the periodic table is for the oxides of the elements to become increasingly acidic as you go from left to right across the periodic table. Fluorine is in Group 7 so, as expected, oxygen difluoride OF_2 is acidic in solution, reacting with water to make hydrofluoric acid and oxygen:

$$OF_2(l) + H_2O(l) \rightarrow 2HF(aq) + O_2(g)$$

• How do the melting points of the elements change as we move from left to right?

The melting points of the elements on the left of each period are quite high, rising to a peak and then falling sharply. For example, in Period 2 they rise from lithium to a peak at carbon, and then fall, the lowest value being that of the noble gas neon. In Period 3 the melting points rise from sodium, though magnesium and aluminium to reach a peak at silicon, with phosphorus, sulphur and chlorine having much lower melting points, and the lowest being that of another noble gas, argon.

• Suggested Films

- Atomic Structure
- The Elements: Potassium
- The Halogens
- The Noble Gases
- Solids, Liquids and Gases

Extension Question

Q6. Why are there two elements in the first period and eight in the second period?

The first period (H, He) corresponds to the filling of the first electron shell, which can only hold 2 electrons. The second period (Li, Be, B, C, N, O, F, Ne) corresponds to the filling of the second electron shell, which holds 8 electrons.

• What pattern is observed in the electrical conductivity of the elements?

The elements found on the left of the period, lithium, sodium, magnesium, are excellent conductors. Those on the right, the non-metals, are poor conductors. In the centre are elements, such as silicon and germanium, which are semi-conductors.

• Suggested Films

- Atomic Structure
- Metallic Bonding
- Carbon: Introduction
- The Elements: Silicon

Section 3: How Do We Explain the Patterns in the Periodic Table?

• What is the scientific basis of the arrangement of the elements in the periodic table?

The elements are arranged in order of their atomic number (the number of protons in the nucleus) 1H, 2He, 3Li, etc. Each element has a unique atomic number, which in turn determines the unique properties of that atom. For example, all fluorine atoms have 9 protons and all sodium atoms have 11 protons, and it is this that makes fluorine so different from sodium.

Once one electron shell is filled the next shell has to be started. For example, neon has 10 electrons, arranged 2,8. Since the second shell is now full, the next element sodium, with 11 electrons, begins a new shell. The electrons are arranged 2,8,1 and thus a new Period begins. As we move from left to right across the period the electron shells are filling up. Atoms with 1, 2 or 3 electrons in the outer shell being mainly metals, and atoms with 5, 6, 7 or 8 electrons in the outer shell being mainly non-metals.

The periods are therefore the horizontal rows in the table, and relate directly to the filling of the shells. They give the table its periodic nature, as elements with similar electron configurations fall into columns called groups. Group 1, the alkali metals (Li, Na, K, Rb, Cs, Fr), all have 1 outer electron, and Group 7, the halogens (F, Cl, Br, I, At), all have 7 outer electrons. The noble gases (Group 0 or Group 8) are particularly significant, as all their electron shells are full, and so they are much less reactive than the other elements. Atoms often combine with other atoms in such a way as to achieve these very stable noble gas configurations of electrons.

• Suggested Films

- Atom Structure: Electron Shells
- Atomic Structure
- The Alkali Metals
- Introduction to Chemical Bonding
- The Halogens
- The Noble Gases
- We Are All Made of Stars

• Why are metals are on the left of the periodic table and non-metals on the right?

To the left of the table, elements in Group 1, 2 and 3 are mostly metals and they have 1, 2 or 3 electrons in the outer shell, respectively. The atoms of these elements can lose their outer electrons relatively easily, and this 'sea' of delocalised electrons can move freely through the structure when a potential difference is applied. They are excellent conductors of electricity.

On the right of the table, in Groups 5 to 8 (Group 8 is also known as Group 0), the atoms have more protons in their nuclei, so the electrostatic attraction between the nuclei and the outer electrons is stronger, making it much harder for the outer electrons to be lost. The electrons cannot move freely, so the elements are poor conductors of electricity. In the centre of the table, in Group 4, carbon, silicon and germanium show properties midway between metals and non-metals. The graphite allotrope of carbon is a good electrical conductor, whereas the diamond allotrope is an insulator. Both silicon and germanium are semi-conductors and are used in computer chips.

• Suggested Films

- Atom Structure: Electron Shells
- Metallic Bonding
- The Elements: Silicon
- Covalent Bonding

Diamond is a poor conductor of electricity

Extension Questions

Q7. An element X is a solid at room temperature, a good conductor of electricity and forms a basic oxide. Where would you expect to find it in the periodic table?

X is a metal element and so would be expected to be somewhere on the left of the periodic table.

Q8. Carbon is a non-metal and yet one allotrope of carbon, graphite, is a good conductor of electricity. Why is this?

Graphite is made of a giant structure of carbon atoms. The atoms are arranged in sheets of interlinked hexagons, in which each carbon atom is covalently bonded to three other carbon atoms. As carbon atoms have four outer electrons, it means that one electron per atom is NOT involved in covalent bonding. It is this electron that is delocalised, making a 'sea' of electrons rather like that in a metal. If a potential difference is applied the electrons can move and make an electric current.

Q9. Why is diamond a very poor conductor of electricity?

Diamond is also made of a giant structure of carbon atoms, but in this case ALL the four outer electrons are involved in making covalent bonds between the carbon atoms. These electrons are therefore fixed in position and cannot move. Since there are no delocalised electrons diamond is a very poor conductor of electricity.

• How do we explain the pattern in the melting points of the elements?

There is strong electrostatic attraction between the metal ions and the ‘sea’ of delocalised electrons – known as metallic bonding – and so a lot of energy is needed to pull the ions away from the electrons, hence the relatively high melting points of these metals.

In the centre of the periods, in Group 4, atoms of carbon, silicon and germanium can form giant structures in which all the atoms are covalently bonded to other atoms. To break up these giant structures requires large amounts of energy and, hence, graphite and diamond (allotropes of carbon), silicon and germanium all have very high melting points.

The elements in Groups 5, 6 and 7 are made of groups of atoms covalently bonded together, known as molecules. Group 0, the noble gases, are made of single atoms. The forces of attraction between these particles – known as intermolecular forces – are quite weak, so not much energy is needed to pull them apart. That is why their melting points are much lower than those of the elements in the first three groups of the period.

• Suggested Films

- Atomic Structure
- Metallic Bonding
- Carbon: Introduction
- Carbon: Synthetic Diamonds
- Covalent Bonding
- The Elements: Silicon
- Intermolecular Forces
- Solids, Liquids and Gases

Extension Question

Q10. Why is oxygen a gas, whereas copper is a solid at room temperature?

Oxygen is made of small O_2 molecules, with only weak intermolecular forces of attraction holding them together in liquid oxygen. It does not take much energy to pull the oxygen molecules apart and change liquid oxygen into a gas. Copper, on the other hand, is a metal with strong electrostatic forces of attraction between positive metal ions and the ‘sea’ of electrons. It requires a lot of energy to pull the copper ions away from the electrons, so copper is a solid at room temperature.

TABLE 04:
Periodic Table of Elements

Periodic Table of the Elements

CHEMISTRY • PERIODIC TABLE • DISCOVERING ELEMENTS

1	2											3	4	5	6	7	0
							H										He
Li	Be											B	C	N	O	F	Ne
Na	Mg											Al	Si	P	S	Cl	Ar
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr
Rb	Sr	Y	Zr	Nb	Mo	Te	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Ti	Pb	Bi	Po	At	Rn
Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg							

Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

• Quizzes

Introduction to the Periodic Table

Basic

• The following are all chemical elements except

A – gold

B – water

C – oxygen

D – hydrogen

• Mendeleev made his Periodic Table in

A – 1968

B – 1908

C – 1896

D – 1869

• It was difficult to find order in the chemical elements because

A – they all looked the same

B – they were all solids

C – they had such different properties

D – they were all gases

• Mendeleev placed the elements in order of their

A – relative atomic mass

B – melting points

C – densities

D – reactivity

Advanced

• Mendeleev wrote on his pack of cards

A – the names of the elements

B – the names of the elements and their relative atomic masses

C – the colours of the elements

D – the names of the elements and their densities

• Mendeleev noticed that the groups had elements with similar

A – chemical properties

B – relative atomic masses

C – names

D – melting points

• Mendeleev left gaps because

A – he was not sure what he was doing

B – he did not have enough cards

C – he could not see what the pattern was

D – he thought more elements would be discovered

• Along the rows, or periods, the elements

A – had increasing relative atomic masses

B – showed increasing reactivity

C – showed similar chemical properties

D – had increasing melting points

Periodic Table and Atomic Structure

Basic

• For a given atom, its atomic number is the number of

A – electrons

B – neutrons

C – protons

D – photons

• Henry Moseley made his suggestion about atomic number in

A – 1703

B – 1913

C – 1923

D – 1932

• Argon atoms have

A – 10 protons

B – 11 protons

C – 17 protons

D – 18 protons

• On Earth, the number of naturally-occurring elements is

A – 54

B – 83

C – 92

D – 100

Advanced

• The element with the smallest atomic number is

A – argon

B – sodium

C – hydrogen

D – helium

• The maximum number of electrons in the second shell is

A – 2

B – 8

C – 18

D – 32

• The number of outer electrons in atoms of the halogen elements is

A – 2

B – 4

C – 7

D – 9

• Elements in the same group have the same number of

A – outer electrons

B – electron shells

C – protons

D – neutrons

• Answers

Introduction to the Periodic Table

Basic

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• Along the rows, or periods, the elements

A – had increasing relative atomic masses

B – showed increasing reactivity

C – showed similar chemical properties

D – had increasing melting points

Periodic Table and Atomic Structure

Basic

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C – 7

D – 9

• Elements in the same group have the same number of

A – outer electrons

B – electron shells

C – protons

D – neutrons