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The Index contains a list of all Help topics available for the Periodic Table. For information on how to use Help, press F1 or choose Using Help from the Help menu.

Periodic Table Commands

To get help with a particular command, choose the appropriate menu.

Option Menu Commands

Set Options

Save Options

Option Menu Commands

The following are the commands available on the option menu.

Set Options

Opens a dialog box showing the currently set options. Make any changes you desire, then choose the OK button. The CANCEL button will exit the dialog without making any changes.

Save Options

Selecting this command will cause Periodic Table to save its current options to the WIN.INI file. These options will be used the next time you start Periodic Table.

Set Options Command

This dialog box has several areas.

Font:

This combo box shows the current font that is used to draw the Periodic Table. By clicking on the dropdown button you will see all available fonts. Highlight the one you want to use and press TAB or click on the dropdown button again.

Colors:

This group of controls allows you to change the colors of the element groups and turn the legend on or off.

The first control is a combo box that contains all the element groups. When you have selected the group you want to change press the SET button. This will bring up the Set Color dialog box. This will allow you to change the color of the group.

The **Show Legend** check box indicates whether the Legend window will be used. If the box is checked the Legend will display, otherwise it will not be displayed.

Detail:

This set of radio buttons selects the type of information to be displayed at the bottom of each element in the table. You may choose Weight, Electrons, Shells, Electronegativity, or Atomic Radius.

Mode:

This set of radio buttons selects the type of information to be displayed in the Legend window and the way the elements are colored. You may choose Groups, Date, or Temperature.

Once the above options have been selected press the OK button to see them. You may press the CANCEL button to abort changing any of the current options.

Using the Legend

The legend will be displayed whenever the Show Legend option is turned on. The default is to display a legend.

The information displayed in the Legend depends on the Mode selected from the Options menu.

Group Mode:

The Legend shows the colors used to represent each group of the Periodic Table.

Date Mode:

The Legend shows the year and the colors used to represent the discovery status of the elements on that date. The date is changed by using the horizontal scroll bar.

Temperature Mode:

The Legend shows the temperature in Kelvin, Celcius, and Fahrenheit and the colors used to represent the physical state of the elements at that temperature. The temperature is changed by using the horizontal scroll bar.

Using the Legend Controls:

To change a group's color, double click on the color box to the left of the group name. The Set Color dialog box will open. You can then change the color.

To remove the Legend you can choose Close from its system menu, double click its system menu icon, press Alt+F4 when the Legend's caption bar is highlighted or toggle the Show Legend checkbox on the Set Options dialog.

To turn the Legend back on choose Show Legend from the Set Options dialog.

Set Color Dialog Box

The Set Color dialog box allows you to change the color of any of the groups of elements. To use the dialog you can use the scroll bars or you can enter a value in the edit control at the end of the scroll bar.

There is a scroll bar for each of the primary colors, Red, Green, and Blue. As you move the slider of a scroll bar to the left the value will increase up to the maximum of 255. You can set each scroll bar independently of the others.

At the right of the dialog is a sample of the color based on the current settings of the scroll bars.

When you have a color you like press the OK button to exit the dialog. The Periodic Table will be redrawn using your new color.

If you don't want to change the color press CANCEL. Any changes made to the current color will be ignored, and you will exit the dialog.

The DEFAULT button will change the color back to the original program defined color. You can then start making changes again or exit the dialog.

Getting Information on an Element

You can get more detail about a particular element shown on the Periodic Table by using the following procedure:

1. Move the mouse pointer over the element you want to know about.
2. Click the left mouse button.

A dialog box will pop-up giving you more information about the element. It will show such things as who discovered the element, where it comes from and what it is used for.

Related Topics

[Element Detail Dialog](#)


Element Detail Dialog

This dialog box is obtained by clicking the left mouse button while the mouse pointer is positioned over an element in the table.


The icon in the upper right corner of the dialog box shows a representation of the element. For example, the element gold (Au) has an icon showing a gold bar. Under the icon is a button that changes depending on the type of element displayed. Press it to display the information on the topic shown on the button.

The box shows additional information on the element. It shows the name, symbol, atomic number, atomic weight, electron shell configuration, subshell being filled, atomic radius, electronegativity, the origin of the name, the discoverer of the element, the discoverer's country, when it was discovered, where the element is usually obtained, and what some of its uses are.

At the bottom of the box are four sets of buttons. Press the **OK** button to close the box. Press the **Help** button to display this help topic.

The pair of buttons on the left allow you to change the element being displayed by moving up  or down

 (moving through the periods) in the Periodic Table.

The pair of buttons on the right allow you to change the element being displayed by moving left  or right

 (moving through the groups) in the Periodic Table.

Exiting

You may exit Periodic Table in one of four ways:

1. Choose Close from the system menu.
2. Press the Alt+F4 key combination.
3. Double click on the system menu icon.
4. Choose Exit from the Options Menu.

Periodic Law

It became evident early in the science of chemistry that certain elements could be grouped together by the similarities of their physical properties. For example, lithium, sodium and potassium have similar properties as do fluorine, chlorine, bromine and iodine.

In 1829, the German chemist Döbereiner, became the first to propose a relationship between atomic weights and the properties of elements. He was able to show that strontium has properties about midway between calcium and barium.

The English chemist Newlands, in 1865, saw a correlation between the atomic weights and the properties of the elements. Mendeleev in Russia and Lothar Meyer in Germany, in 1870, working independently and unaware of Newlands, outlined this relationship in detail.

This led to a statement that the properties of the elements are periodic functions of their atomic weights. The below series illustrates the point:

Li	Be	B	C	N	O	F	Ne
6.94	9.01	10.81	12.01	14.01	16.00	18.00	20.18
Na	Mg	Al	Si	P	S	Cl	Ar
22.00	24.31	26.98	28.09	30.97	32.06	35.45	39.95

In each row elements differ quite a bit in properties. There is, however, an interesting gradation of properties as atomic weight increases. Note that elements in a vertical column have similar properties. For example, lithium is similar to sodium.

Moseley showed that atomic number rather than atomic weights are the determining factor of chemical properties. If you check the Periodic Table you can see that argon (At. Wt. 39.948) precedes potassium (At. Wt. 39.098). There are other reversals in the table as well. This is due to the abundance of heavier isotopes of some elements. This affects the atomic weight which is an average.

The formal statement of the Periodic Law is: **The properties of the elements are periodic functions of their atomic numbers.**

The atomic weight of an element is the average relative weight of an element referred to carbon 12 as a standard.

Atomic number is an experimentally determined number characteristic of a chemical element that represents the number of protons in the nucleus which in a neutral atom equals the number of electrons outside the nucleus.

Two or more species of atoms having the same atomic number (number of protons) but with differing atomic mass number (number of protons and neutrons). In other words they will differ in the number of neutrons in the nucleus. Isotopes have nearly the same chemical properties but may differ in physical properties.

Electron Shells

As atomic number increases so too does the number of electrons in a neutral atom of the element. The valence electrons are largely responsible for its chemical behavior. If elements having the same number of valence electrons are grouped together the elements in each group or family will have similar chemical properties.

The **first short period** contains only two elements. These elements fill the K shell. It ends with helium which contains a full K shell consisting of two electrons.

The **second short period** contains eight elements beginning with lithium and ending with neon. Neon has a complete L shell of eight electrons.

The **third short period** contains eight elements beginning with sodium. It ends with argon, which contains eight electrons in the M shell.

The **fourth period** contains eighteen elements beginning with potassium. It includes a series of elements from scandium through copper. These elements are known as **transition elements**.

They are building the M shell from eight to eighteen electrons. Before this series begins though two electrons enter the outermost N shell. After the series the N shell then builds up to eight electrons ending with krypton.

The **fifth period** is just like the fourth period in which the O shell takes two electrons then the N shell builds to eighteen electrons before the O shell continues out to eight electrons ending with xenon.

The **sixth period** contains 32 elements. It too has a transition series beginning with lanthanum and ending with gold, in which the O shell fills to eighteen electrons.

Note however that this series is split after lanthanum with a series of fourteen elements. In these elements the N shell (third from the outside) is filling from eighteen to 32 electrons. These are referred to as the **inner transition series, lanthanum series** or the **rare earth elements**. **It begins with cerium and ends with lutetium.**

Following lutetium the transition series from hafnium to gold completes by building the O shell to eighteen electrons. The outer P shell then continues to build to eight electrons ending with radon.

The **seventh period** is incomplete. The first two elements, francium and radium fill the Q shell with two electrons. Then the P and O shells fill in the same way as the sixth period. This produces the **actinide series**.

Related Topics

[Electron Subshells](#)

The electrons in the outermost shell of an atom are called its valence electrons.

Each horizontal row in the Periodic Table is called a period. Each period begins filling a new electron shell.

Electron Subshells

Each electron shells, designated by the letters K, L, M, N, O, P, and Q or the numbers 1, 2, 3, 4, 5, 6, and 7, are subdivided further into **subshells**.

The first principal shell, the K shell, has one subshell designated as the $1s$ shell. The second principal shell has two subshells designated as $2s$ and $2p$. The third shell has three subshells designated $3s$, $3p$, and $3d$. The fourth shell has four subshells, $4s$, $4p$, $4d$, and $4f$.

The subshells are further divided into **orbitals**, each can contain a maximum of two electrons. Therefore an s subshell, which is made up of one orbital, can contain a maximum of two electrons. A p subshell has three orbitals and can contain six electrons. A d subshell has five orbitals and can contain ten electrons. An f subshell has seven orbitals and can contain fourteen electrons.

On the average, an s electron will approach the nucleus more closely than a p of the same principal shell, a p electron more closely than a d electron, and a d electron more closely than an f electron.

The notation used in describing electron structures consists of a principal shell number in front of the subshell letter and a superscript to designate the number of electrons in that particular subshell. For example $3d^4$ indicates that there are four electrons in subshell d of the M shell.

Related Topics

[Shapes of Electron Orbitals](#)

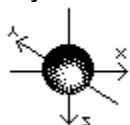
Electron Orbital Shapes

The orbitals occupied by electrons in different subshells differ from each other with regard to their size, shape, and orientation in space. The larger the number of the principal shell, the greater the volume of the corresponding orbital.

For example, a 2s electron has an orbital of greater volume than a 1s electron. However, the orbitals of the higher principal shells overlap those of lower shells. This means that the outer electrons penetrate the regions occupied by inner electrons.

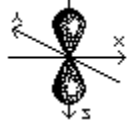
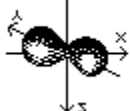
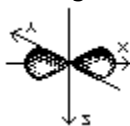
The shape of an s electron orbital is spherical while that of a p electron orbital is dumbbell shaped.

The number of possible orientations of atomic orbitals in space depends to an extent on the shape of the orbital. An s orbital still presents the same appearance to an outside observer since its shape is that of a sphere. Therefore an s orbital is said to be spherically symmetrical and without directional characteristics.



s orbitals

It is a bit more complex with p orbitals which occur in sets of three. It is not possible to determine the direction of any one orbital in a given set. However, the axes along which the three p orbitals lie are mutually at right angles to each other. The three p orbitals are designated as p_x , p_y and p_z to emphasize their directional nature.

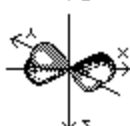
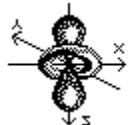


p_x

p_y

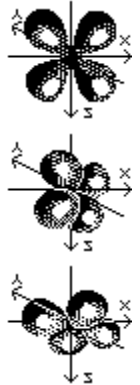
p_z

The five d orbitals consist of lobe-shaped regions arranged in space as shown below. The drawings are oriented so that the lobes of the orbitals designated d_{z^2} and $d_{x^2-y^2}$ lie along the axes and the lobes of the d_{xz} , d_{yz} , and d_{xy} lie in between the axes.



d_{z^2}

$d_{x^2-y^2}$



dxz

dyz

dxy

The seven *f* orbitals are not well understood enough to have agreement on their orientation in space. Their directional characteristics are too complex to be shown in this help file.

Rare Earth Elements

These elements got their name from the fact that chemists first isolated them in their oxide forms. These oxides somewhat resemble calcium, magnesium and aluminum oxides, sometimes called common earths.

The rare earths all have three valence electrons. Because of this they all have similar properties.

Related Topics

[Metals](#)

[Transition Metals](#)

[Alkali Metals](#)

[Alkali Earth Metals](#)

Transition Metal Elements

This group of metals are distinguished from other metals not by their physical properties, but by their electronic structure. The transition metals have their valence electrons in more than one shell. Regular, or representative, metals have their valence electrons in only one shell.

The transition metals include the elements of the groups IIIA through IIB in each of the long periods of the periodic table.

The transition metals are noted for their variability in oxidation state. This is attributed to the presence of valence electrons in more than one shell. Whereas, the representative metals have only one or two oxidation states, since they have valence electrons in only one shell.

Thus, manganese has two electrons in its outside shell and five electrons in the next shell down, and exhibits oxidation states of +1, +2, +3, +4, +5, +6, and +7.

They are also characterized by the fact that well into the series, going from left to right, the properties of succeeding metals do not differ greatly from preceding ones. This is attributed to the fact that, generally, succeeding elements differ in electronic structure by one electron in the next to the outer valence shell rather than the outer shell.

In contrast, those of succeeding representative metals in a period differ extensively, since they differ by one electron in the outer valence shell.

Related Topics

[Metals](#)

[Akali Metals](#)

[Alkali Earth Metals](#)

[Rare Earth Metals](#)

Alkali Metal Elements

Alkali metal is a term that refers to six elements: lithium, sodium, potassium, rubidium, cesium, and francium. These elements make up group IA of the periodic table of elements. They all form singly-charged positive ions, and are extremely reactive. They react violently with water, forming hydroxides and releasing hydrogen gas and heat. Cesium and francium are the most reactive and lithium is the least.

Related Topics

[Metals](#)

[Alkali Earth Metals](#)

[Transition Metals](#)

[Rare Earth Metals](#)

Alkali Earth Metal Elements

Alkali earth metal is a term that refers to six elements: beryllium, magnesium, calcium, strontium, barium, and radium. These elements make up group IIA of the periodic table of elements. They all exhibit a single oxidation state, +2. They are all light and very reactive. Barium and radium are the most reactive and beryllium is the least.

The term "earth" was formerly used by chemists to denote slightly soluble metal oxides. The oxides of barium, strontium, and calcium resemble alumina (Al_2O_3), a typical "earth", but form alkaline mixtures with water. For this reason barium, strontium, and calcium were called alkaline earth metals. This name has now been extended to include all of the elements of group IIA.

Related Topics

[Metals](#)

[Alkali Metals](#)

[Transition Metals](#)

[Rare Earth Metals](#)

Metal Elements

Chemists use a test called an electrolysis test to determine if an element is metallic or non-metallic. This test consists of dissolving the element in acid and running an electric current through the solution. If the element is metallic, the atoms of the element will show a positive charge. That is they will seek the negative pole.

Formally it is stated as **"those elements which, when in solution in a pure state, carry a positive charge and seek the negative pole. in an electric cell"**.

Properties of metals that are familiar to everyone include:

1. They are good conductors of heat and electricity
2. They are usually opaque to light
3. They have metallic luster

The difference in chemical properties between metals and non-metals lies mainly in the fact that atoms of non-metals will readily fill their valence shells by sharing electrons with or transferring electrons to other atoms.

For example, chlorine, a typical non-metal, will readily enter into combination with another atom by adding one electron to an outer shell of seven: gaining an electron from or sharing a pair with the other atom. Whereas, an atom of say sodium, a typical metal, enters into combination by the loss of its single valence electron.

Metallic characteristics of the elements decrease and non-metallic characteristics increase with the increase of valence electrons. Also metallic characteristics increase with the number of electron shells. Therefore there is no sharp dividing line between the metals and the non-metals.

There are elements that are definitely on the border between metal and non-metal. These elements will show hybrid behavior. They are sometimes called **metalloids**. Many of them have been called semiconductors. This group of elements include boron, silicon, germanium, arsenic, antimony, tellurium, and astatine.

Their physical properties generally include:

1. They are good conductors of heat and electricity.
2. They are malleable and ductile in their solid state.
3. They show metallic luster.
4. They are opaque.
5. They have high density.
6. They are solids (except mercury)
7. They have crystal structure in which each atom is surrounded by eight to twelve near neighbors

Their chemical properties generally are:

1. They have one to four valence electrons.
2. They have low ionization potentials; they readily lose electrons.
3. They are good reducing agents.
4. They have hydroxides which are bases or amphoteric.
5. They are electropositive.

Of the 110 elements now known, only 17 show primarily non-metallic characteristics, 7 others are metalloids, and 86 may be classed as metals.

Related Topics

[Non-Metals](#)

[Transition Metals](#)

[Alkali Metals](#)

[Alkali Earth Metals](#)

[Rare Earth Metals](#)

[Halogen Elements](#)

[Noble Gas](#)

Halogen Elements

The halogens are the elements astatine, fluorine, chlorine, bromine, and iodine. They are non-metals, and make up part of the seventh group in the periodic table.

The halogens all have a strong unpleasant odor and will burn flesh. They do not dissolve well in water.

The five elements are strongly electronegative. They are oxidizing agents, with fluorine being the strongest and astatine being the weakest. They react with most metals and many non-metals.

As the atomic structure of the halogens becomes more complex with increasing atomic weight there is a gradation in physical properties. For example: Fluorine is a pale yellow-green gas of low density. Chlorine is a greenish-yellow gas 1.892 times as dense as fluorine. Bromine is a deep reddish-brown liquid which is three times as dense as water. Iodine is a grayish-black crystalline solid with a metallic appearance. And astatine is a solid with properties which indicate that it is somewhat metallic in character.

Related Topics

[Non-Metal Elements](#)

[Noble Gas Elements](#)

Noble Gas Elements

Noble gas refers to any element of the group of six elements in group VIII of the periodic table. They are argon, helium, krypton, neon, radon, and xenon. Unlike most elements, the noble gases are monoatomic. The atoms have stable configurations of electrons. Therefore under normal conditions they do not form compounds with other elements.

They were generally called inert gases until about 1962 when xenon tetrafluoride, XeF₄, was produced in the laboratory. **This was the first report of a stable compound of a noble gas with another single element.**

Related Topics

[Non-Metal Elements](#)

[Halogen Elements](#)

Non-Metal Elements

Non-metals are basically defined as elements that are not metals.

Their physical properties generally include:

1. They are poor conductors.
2. They are brittle, not ductile in their solid state.
3. They show no metallic luster.
4. They may be transparent or translucent.
5. They have low density.
6. They are gases, liquids, or solids.
7. They form molecules which consists of atoms covalently bonded; the noble gases are monoatomic.

Their chemical properties are generally:

1. They usually have four to eight **valence electrons**.
2. They have high electron affinities. (except the noble gases)
3. They are good oxidizing agents. (except the noble gases)
4. They have hydroxides which are acidic. (except the noble gases)
5. They are electronegative.

Related Topics

Metals

Transition Metals

Alkali Metals

Alkali Earth Metals

Rare Earth Metals

Halogen Elements

Noble Gas

A substance is ductile if it can be hammered out thin or drawn into a wire.

The ability to form a bond in which two atoms share a pair of electrons.

The tendency to gain electrons; nonmetallic; acid.

The tendency to lose electrons; metallic; basic.

Any compound consisting of an element or radical combined with one or more hydroxyl radicals (OH - one oxygen atom combined with one hydrogen atom).

Change the atoms or ions of an element to a higher positive valence by the loss of electrons. An oxidizing agent is an element that can remove electrons from another element.

Change the atoms or ions of an element to a lower positive valence by the gain of electrons. A reducing agent is an element that can add electrons to another element.

A compound that reacts as a weak acid or a weak base.

The state of an element in a compound with respect to the number of electrons it has lost or gained, expressed as a positive or negative number indicating the ionic charge of the atom and equal to its valence.

Halogen means salt producer. Sea salts are compounds of halogens with metals.

The term earth was formerly used by chemists to denote slightly soluble metal oxides.

Monoatomic means it occurs as single atoms rather than in pairs or combination.