



高等院校“十三五”系列规划教材

# 化学英语

主 编 王 帅  
副主编 李迎春

Chief Editor Wang Shuai  
Associate Editor Li Yingchun

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专业英语——行业英语



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# 化 学 英 语

## CHEMISTRY ENGLISH

主 编 王 帅  
副主编 李迎春  
参 编 马晓利 赵志西

Chief Editor  
Wang Shuai  
College of Chemistry and Chemical Engineering,  
Key Laboratory of Pollution Monitoring and Control  
Xinjiang Normal University

Associate Editor  
Li Yingchun  
School of Pharmacy,  
Shihezi University

Ma Xiaoli  
Key Laboratory of Pollution Monitoring and Control  
Xinjiang Normal University

Zhao Zhixi  
Key Laboratory of Pollution Monitoring and Control  
Xinjiang Normal University

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# Preface

英语作为世界上主要的国际通用语言,使用范围非常广泛。全世界科技出版物70%以上用英语发表,绝大部分的国际会议是以英语为通用语言,因此,英语对理工科学生的重要性不言而喻。随着高校国际化程度的不断提高,学术交流活动的日益频繁,专业英语的重要性愈发凸显;专业化学英语教学对于化学专业本科生和研究生掌握本领域的科学知识和技能有着特殊的意义,是其将来进行科学研究的基石。

然而,当前化学专业英语教材普遍存在着偏重英语忽视专业的现象,专业英语授课中不得不花大量时间介绍语法、时态等基础概念,这与大学英语相重叠,浪费了宝贵的教学课时;也有少数教材侧重专业介绍,以英文科技论文和学科介绍为主,但是对于大多数本科生而言过于专业,阅读学习难度太大,教师教学也困难重重。

由于新疆地区特有的多民族多语言现状,学生英语成绩普遍较差,部分民族同学甚至没有学过英语,造成班级中学生的英语水平差距明显。目前已出版的偏重化学专业的专业英语教材对于学生英文要求的起点一般都比较高,造成授课的困难,这是目前专业英语教学面临的主要问题;而解决这一问题的最好途径就是根据学生实际情况编辑适合本科生的专业英语书籍。

本教材在教学内容上侧重专业知识介绍,根据三年级本科生已有一定专业知识的情况,介绍元素周期表及每一个元素的性质和特点。本教材以元素周期表介绍为第一章开始,一至八章分别介绍八个主族元素的历史、词源、性质及特点,九、十章分别介绍d区和f区元素,附录部分收录了环境化学、化工和生物化学等专业词汇,起到专业词汇词典作用。内容上突出专业英语课与普通英语课的区别,重点强调其专业性,与所学化学专业课程内容结合,增强化学专业语境;将学生已学的专业知识用英语再学习、再巩固,降低了英语基础要求,增强了学生的学习能动性。在教学中通过介绍前沿科学、历史及人物故事、词根来源等,极大地增强了学生的学习积极性,对专业名词的记忆有很好的帮助;同时,此教材主要以英文编著,帮助学生克服对英文书籍的恐惧,增强其阅读英文教材及文献的自信心。

本教材的编辑得到了新疆师范大学和石河子大学校领导、教务处、化学化工学院等领导的大力指导和帮助。北京师范大学出版社易新老师对本书的编辑出版给予了积极的支持,在此一并表示感谢。

由于执教化学专业英语时间不长,加之作者本身水平和视野有限,本书难免存在一些纰漏甚至错误,诚恳地希望读者予以指正。

王帅 李迎春

2014年12月于新疆师范大学

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# Chapter 0 Periodic Table

Layout, History, Grouping methods, Periodic trends, Atomic structure  
元素周期表，通往化学世界的大门！

Group→	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
↓Period																		
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117 Uus	118 Uuo
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

A periodic table is a tabular display of the chemical elements, organized on the basis of their atomic numbers, electron configurations, and recurring chemical properties. Elements are presented in order of increasing atomic number (number of protons). The colors represent different categories. The table can also be deconstructed into four rectangular blocks; the s-block to the left, the p-block to the right, the d-block in the middle, and the f-block below that.

## 0.1 Layout

The standard form of table comprises an 18 × 7 grid or main body of elements, positioned above a smaller double row of elements. In the standard periodic table, the elements are listed in order of increasing atomic number (the number of protons in the nucleus of an atom). A new row (period) is started when a new electron shell has its first electron. Columns (groups) are determined by the electron configuration of the atom; elements with the same

number of electrons in a particular subshell fall into the same columns (e. g. oxygen and selenium are in the same column because they both have four electrons in the outermost *p*-subshell).

Since, by definition, a periodic table incorporates recurring trends, any such table can be used to derive relationships between the properties of the elements and predict the properties of new, yet to be discovered or synthesized, elements.

All elements from atomic number 1 (hydrogen) to 118 (ununoctium) have been discovered or synthesized. Of these, all up to and including californium (atomic number 98) exist naturally; the rest have only been synthesized in laboratories. Production of elements beyond ununoctium (118) is being pursued, with the question of how the periodic table may need to be modified to accommodate any such additions being a matter of ongoing debate. Numerous synthetic radionuclides of naturally occurring elements have also been produced in laboratories. As of 2012, the periodic table contains 118 confirmed chemical elements. Of these elements, 114 have been officially recognized and named by the International Union of Pure and Applied Chemistry (IUPAC).

0. 2 History

Russian chemistry professor Dmitri Mendeleev and German chemist Julius Lothar Meyer independently published their periodic tables in 1869 and 1870, respectively. Mendeleev's table was his first published version; that of Meyer was an expanded version of his (Meyer's) table of 1864. They both constructed their tables by listing the elements in rows or columns in order of atomic weight and starting a new row or column when the characteristics of the elements began to repeat.

ОПЫТЪ СИСТЕМЫ ЭЛЕМЕНТОВЪ. ОСНОВАННОЙ НА ИХЪ АТОМНОМЪ ВѢСѢ И ХИМИЧЕСКОМЪ СХОДСТВѢ.					
		Ti=50	Zr=90	? = 180.	
		V = 51	Nb=94	Ta= 182.	
		Cr=52	Mo=96	W= 186.	
		Mn=55	Rh= 104,4	Pt= 197,1.	
		Fe=56	Rn= 104,4	Ir= 198.	
		Ni=Co=59	Pl= 106,6	Os= 199.	
H= 1			Cu=63,1	Ag= 108	Hg=200.
	Be=9,1	Mg=24	Zn=65,2	Cd= 112	
	B= 11	Al= 27,1	? =68	Ur= 116	Lu= 197?
	C= 12	Si=28	? =70	Sn= 118	
	N= 14	P=31	As=75	Sb= 122	Bi=210?
	O= 16	S=32	Se=79,1	Te= 128?	
	F= 19	Cl=35,6	Br=80	I= 127	
Li=7	Na=23	K=39	Rb=85,1	Cs= 133	Tl=204.
		Ca=40	Sr=87,1	Ba= 137	Pb=207.
		? =45	Ce=92		
		? Er=56	La=94		
		? Yl=60	Di=95		
		? In=75,6	Th= 118?		

Д. Мелдадвомъ

Fig. 0. 1 Mendeleev's 1869 periodic table



The recognition and acceptance afforded Mendeleev's table came from two decisions he made. The first was to leave gaps in the table when it seemed that the corresponding element had not yet been discovered, such as gallium and germanium. The second decision was to occasionally ignore the order suggested by the atomic weights and switch adjacent elements, such as tellurium and iodine, to better classify them into chemical families. With the development of theories of atomic structure, it became apparent that Mendeleev had unintentionally listed the elements in order of increasing atomic number or nuclear charge.

## 0.3 Grouping methods

### 0.3.1 Groups

A *group* or *family* is a vertical column in the periodic table. Groups usually have more significant periodic trends than periods and blocks, explained below. Modern quantum mechanical theories of atomic structure explain group trends by proposing that elements within the same group generally have the same electron configurations in their valence shell. Consequently, elements in the same group tend to have a shared chemistry and exhibit a clear trend in properties with increasing atomic number. However in some parts of the periodic table, such as the d-block and the f-block, horizontal similarities can be as important as, or more pronounced than, vertical similarities.

Elements in the same group tend to show patterns in atomic radius, ionization energy, and electronegativity. From top to bottom in a group, the atomic radii of the elements increase. Since there are more filled energy levels, valence electrons are found farther from the nucleus. From the top, each successive element has a lower ionization energy because it is easier to remove an electron since the atoms are less tightly bound. Similarly, a group has a top to bottom decrease in electronegativity due to an increasing distance between valence electrons and the nucleus. There are exceptions to these trends, however, an example of which occurs in group 11 where electronegativity increases farther down the group.

### 0.3.2 Periods

A *period* is a horizontal row in the periodic table. Although groups generally have more significant periodic trends, there are regions where horizontal trends are more significant than vertical group trends, such as the f-block, where the lanthanides and actinides form two substantial horizontal series of elements.

Elements in the same period show trends in atomic radius, ionization energy, electron affinity, and electronegativity. Moving left to right across a period, atomic radius usually decreases. This occurs because each

successive element has an added proton and electron which causes the electron to be drawn closer to the nucleus. This decrease in atomic radius also causes the ionization energy to increase when moving from left to right across a period. The more tightly bound an element is, the more energy is required to remove an electron. Electronegativity increases in the same manner as ionization energy because of the pull exerted on the electrons by the nucleus. Electron affinity also shows a slight trend across a period. Metals (left side of a period) generally have a lower electron affinity than nonmetals (right side of a period), with the exception of the noble gases.

### 0.3.3 Blocks

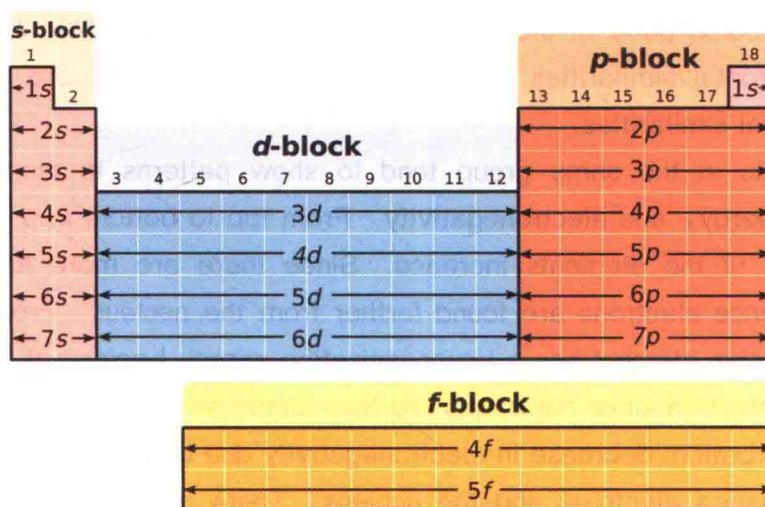


Fig. 0.2 Diagram of the periodic table, highlighting the different blocks

Because of the importance of the outermost electron shell, the different regions of the periodic table are sometimes referred to as *blocks*, named according to the subshell in which the “last” electron resides.

The s-block comprises the first two groups (alkali metals and alkaline earth metals) as well as hydrogen and helium.

The p-block comprises the last six groups which are groups 13 to 18 in IUPAC and contains, among other elements, all of the metalloids.

The d-block comprises groups 3 to 12 in IUPAC and contains all of the transition metals.

The f-block, usually offset below the rest of the periodic table, comprises the lanthanides and actinides.

## 0. 4 Periodic trends

### 0. 4. 1 Electron configuration

The electron configuration or organisation of electrons orbiting neutral atoms shows a recurring pattern or periodicity. The electrons occupy a series of electron shells (numbered shell 1, shell 2, and so on). Each shell consists of one or more subshells (named s, p, d, f and g). As atomic number increases, electrons progressively fill these shells and subshells more or less according to the

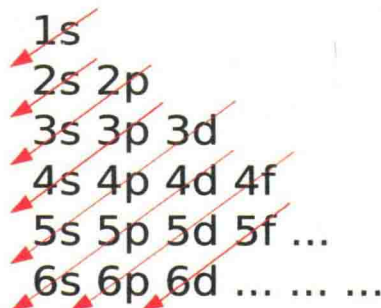


Fig. 0. 3 Madelung rule

Madelung rule or energy ordering rule, as shown in the diagram below. The electron configuration for neon, for example, is  $1s^2 2s^2 2p^6$ . With an atomic number of ten, neon has two electrons in the first shell, and eight electrons in the second shell — two in the s subshell and six in the p subshell. In periodic table terms, the first time an electron occupies a new shell corresponds to the start of each new period, these positions being occupied by hydrogen and the alkali metals.

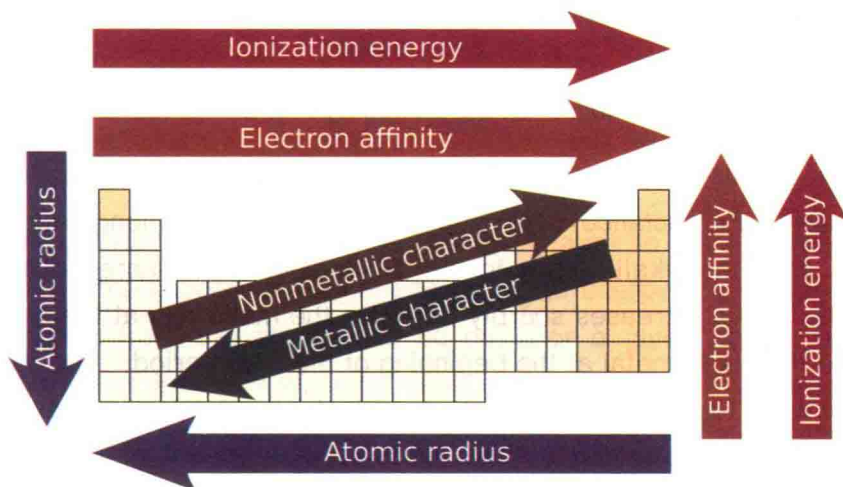


Fig. 0. 4 Periodic table trends. Arrows point at increase

Since the properties of an element are mostly determined by its electron



configuration, the properties of the elements likewise show recurring patterns or periodic behaviour, some examples of which are shown in the diagram below. It is this periodicity of properties, manifestations of which were noticed well before the underlying theory was developed, that led to the establishment of the periodic law (the properties of the elements recur at varying intervals) and the formulation of the first periodic tables.

#### 0. 4. 2 Atomic radius

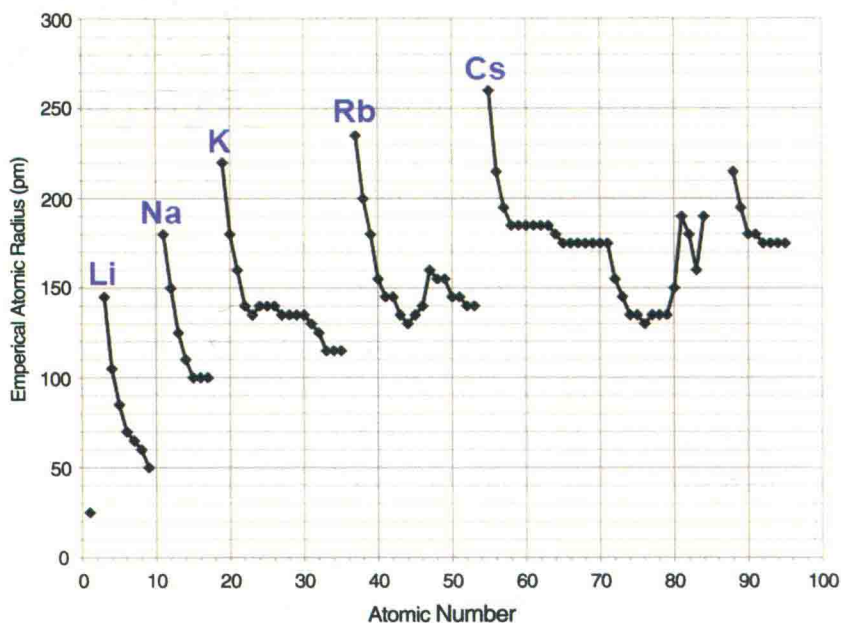


Fig. 0. 5 Atomic number plotted against atomic radius

Atomic radius vary in a predictable and explainable manner across the periodic table. For instance, the radii generally decrease along each period of the table, from the alkali metals to the noble gases; and increase down each group. The radius increases sharply between the noble gas at the end of each period and the alkali metal at the beginning of the next period. These trends of the atomic radii (and of various other chemical and physical properties of the elements) can be explained by the electron shell theory of the atom; they provided important evidence for the development and confirmation of quantum theory.

The electrons in the 4f-subshell, which is progressively filled from cerium

( $Z = 58$ ) to lutetium ( $Z = 71$ ), are not particularly effective at shielding the increasing nuclear charge from the sub-shells further out. The elements immediately following the lanthanides have atomic radii which are smaller than would be expected and which are almost identical to the atomic radii of the elements immediately above them. Hence hafnium has virtually the same atomic radius (and chemistry) as zirconium, and tantalum has an atomic radius similar to niobium, and so forth. This is known as the lanthanide contraction. The effect of the lanthanide contraction is noticeable up to platinum ( $Z = 78$ ), after which it is masked by a relativistic effect known as the inert pair effect. The d-block contraction, which is a similar effect between the d-block and p-block, is less pronounced than the lanthanide contraction but arises from a similar cause.

### 0. 4. 3 Ionization energy

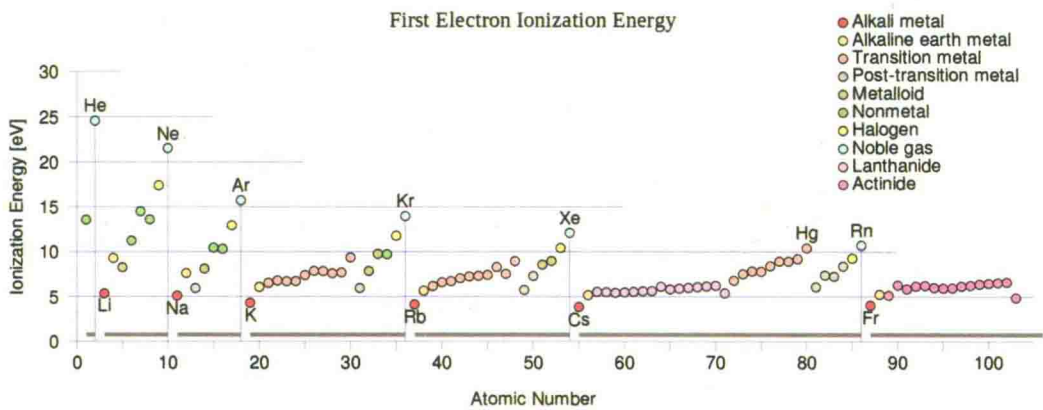


Fig. 0. 6 Ionization energy vs atomic radius

The first ionization energy is the energy it takes to remove one electron from an atom, the second ionization energy is the energy it takes to remove a second electron from the atom, and so on. For a given atom, successive ionization energies increase with the degree of ionization. For magnesium as an example, the first ionization energy is 738 kJ/mol and the second is 1450 kJ/mol. Electrons in the closer orbitals experience greater forces of electrostatic attraction; thus, their removal requires increasingly more energy. Ionization energy becomes greater up and to the right of the periodic table.

Large jumps in the successive molar ionization energies occur when removing an electron from a noble gas (complete electron shell) configuration. For magnesium again, the first two molar ionization energies of magnesium given above correspond to removing the two 3s electrons, and the third ionization energy is a much larger 7730 kJ/mol, for the removal of a 2p electron from the very stable neon-like configuration of  $\text{Mg}^{2+}$ . Similar jumps occur in the ionization energies of other third-row atoms.

#### 0. 4. 4 Electronegativity

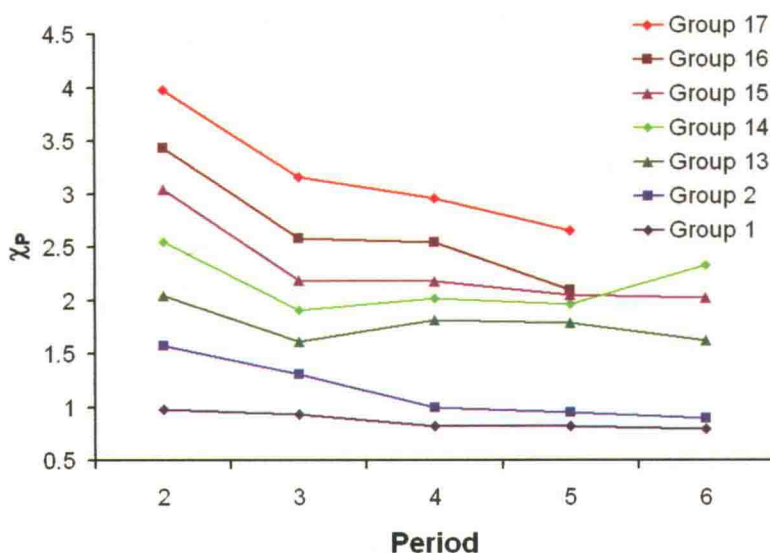


Fig. 0. 7 Electronegativity increasing with growing number of selected groups

Electronegativity is the tendency of an atom to attract electrons. An atom's electronegativity is affected by both its atomic number and the distance between the valence electrons and the nucleus. The higher its electronegativity, the more an element attracts electrons. It was first proposed by Linus Pauling in 1932. In general, electronegativity increases on passing from left to right along a period, and decreases on descending a group. Hence, fluorine is the most electronegative of the elements, while caesium is the least, at least of those elements for which substantial data is available.

There are some exceptions to this general rule. Gallium and germanium have higher electronegativities than aluminium and silicon respectively because

of the d-block contraction. Elements of the fourth period immediately after the first row of the transition metals have unusually small atomic radii because the 3d-electrons are not effective at shielding the increased nuclear charge, and smaller atomic size correlates with higher electronegativity. The anomalously high electronegativity of lead, particularly when compared to thallium and bismuth, appears to be an artifact of data selection (and data availability)—methods of calculation other than the Pauling method show the normal periodic trends for these elements.

0. 4. 5 Electron affinity

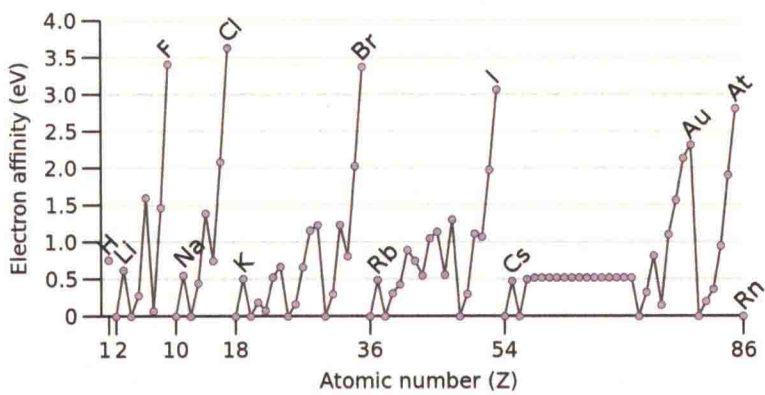


Fig. 0. 8 Dependence of electron affinity on atomic number

The electron affinity of an atom is the amount of energy released when an electron is added to a neutral atom to form a negative ion. Although electron affinity varies greatly, some patterns emerge. Generally, nonmetals have more positive electron affinity values than metals. Chlorine most strongly attracts an extra electron. The electron affinities of the noble gases have not been measured conclusively, so they may or may not have slightly negative values.

Electron affinity generally increases across a period. This is caused by the filling of the valence shell of the atom; a group 17 atom releases more energy than a group 1 atom on gaining an electron because it obtains a filled valence shell and is therefore more stable.

A trend of decreasing electron affinity going down groups would be expected. The additional electron will be entering an orbital farther away from



the nucleus. As such this electron would be less attracted to the nucleus and would release less energy when added. However, in going down a group, around one-third of elements are anomalous, with heavier elements having higher electron affinities than their next lighter congeners. Largely, this is due to the poor shielding by d and f electrons. A uniform decrease in electron affinity only applies to group 1 atoms.

#### 0. 4. 6 Metallic character

The lower the values of ionization energy, electronegativity and electron affinity the more metallic character the element has. Conversely, nonmetallic character increases with higher values of these properties. Given the periodic trends of these three properties, metallic character tends to decrease going across a period and, with some irregularities (mostly) due to poor screening of the nucleus by d and f electrons, and relativistic effects, tends to increase going down a group. Thus, the most metallic elements (such as caesium and francium) are found at the bottom left of traditional periodic tables and the most nonmetallic elements (oxygen, fluorine, chlorine) at the top right. The combination of horizontal and vertical trends in metallic character explains the stair-shaped dividing line between metals and nonmetals found on some periodic tables, and the practice of sometimes categorizing several elements adjacent to that line, or elements adjacent to those elements, as metalloids.