

Year 11 – Chemistry

The Periodic Table

History of the Periodic Table of the Elements

Introduction

The periodic table was developed by chemists as a way of systematically displaying their knowledge of the elements in a concise format. In the periodic table elements are arranged in increasing order of atomic number. The history of the development of the periodic table stretches more than 2000 years and is continually being updated as scientists create new elements.

Historical Development

Reference: 1:2:3:4:5:6

Aristotle (approximately 300BC)

The first attempt to classify elements began with Aristotle who proposed that all matter consisted of a mixture of the four elements; earth, air, fire, and water.

Antoine Lavoisier

It was however not until the works of Antoine Lavoisier before any significant attempt at classifying elements was achieved. Lavoisier divided up the known elements (at the time 33) into four separate groups, this distinguished between some metals and non-metals.

Table 1.1 Lavoisier's 'Table of simple substances'

First Group	Second group	Third Group			Fourth Group
Light	Sulfur	Antimony	Iron	Platinum	Lime
Heat	Phosphorus	Arsenic	Lead	Silver	Magnesia
Oxygen	Carbon	Bismuth	Manganese	Tin	Baryta
Nitrogen	Murium (C)	Cobalt	Mercury	Tungsten	Alumina
Hydrogen	Fluorum (F)	Copper	Molybdenum	Zinc	Silica
	Boracum (B)	Gold	Nickel		

Note: Murium is now known as carbon; Fluorum is known as Fluorine; and Boracum is known as Boron.

Jakob Berzelius

The works of Jakob Berzelius in the early 1800's also aided the development of the periodic table and his interest in a variety of compounds led to his discovery of a number of new elements, including cerium, selenium, and thorium. Berzelius determined the atomic weights of virtually all the elements identified at the time. Since he was dealing with so many elements in so many compounds he found a simple and logical system of symbols.

Johann Döbereiner (1829)

Johann Döbereiner suggested in 1829 the idea that elements could be classified according to their chemical properties. Döbereiner identified groups of elements, with similar chemical and physical properties, which he called 'triads'.

Table 1.2 Döbereiner's "Triads"

Chlorine	Sulfur	Calcium	Lithium
Bromine	Selenium	Strontium	Sodium
Iodine	Tellurium	Barium	Potassium

Alexandre Beguyer de Chancourtois (1862)

In 1862, Chancourtois developed a helix (or Telluric Screw) on a vertical cylinder on which he arranged the elements in order of increasing atomic weight down the cylinder. The cylinder had a circumference (or horizontal axis) of 16 equal parts, and the atomic weights written up on the vertical axis.

John Newlands (1865)

In 1865 a chemist named John Newlands constructed the first periodic table of elements (60 known at that time) whereby elements were arranged in order of increasing atomic weight. He then proposed the "law of octaves", which basically states that for any given element, an element eight places away is similar in

properties. However it has been shown that his law only applies up to the element calcium, after which his law fails.

Table 1.3 Newlands's list of elements, 1865.

No.	No.	No.	No.	No.	No.	No.	No.
H	1	F	8	Cl	15	Co,	22
						Ni	
Li	2	Na	9	K	16	Cu	23
G	3	Mg	10	Ca	17	Zn	24
B	4	Al	11	Cr	18	Y	25
						Ce,	32
C	5	Si	12	Ti	19	In	26
N	6	P	13	Mn	20	As	27
						Di,	34
O	7	S	14	Fe	21	Se	28
						Rh,	35
						Ru	

Lothar Meyer (1969)

Lothar Meyer in 1969, independently and concurrently to Mendeleev, produced a periodic table of the 56 known elements. His table also was based on the arrangement in increasing order of atomic weight. He also separately devised the periodic law.

Dmitri Mendeleev (1969)

In 1869 a Russian chemist named Dmitri Mendeleev, produced a periodic table which arranged the elements in increasing order of atomic weight, and also placed elements with similar properties in the same column or group. He made corrections to the table whenever atomic weights did not agree with properties and by identifying large gaps between atomic weights he was able to predict certain elements (and their properties) before they were discovered. For example, he left gaps for the elements Gallium and Germanium which had not been discovered at the time, he was even able to predict the properties which were extremely accurate (see table 1.4 for details). Mendeleev also proposed the periodic law, which states that: the chemical and physical properties of an element are periodic function of the atomic weight. We now know that it is the atomic number that is the number of protons in the nucleus, which determines these properties; however the principal is basically the same.

Table 1.5 Mendeleev's 1871 Periodic Table

Row	Group 1	Group 2	Group 3	Group 4	Group 5	Group 6	Group 7	Group 8
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		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		Cd=112	In=113	Sn=118	Sb=122	Te=125	I=127	
8	Cs=133	Ba=137	Di=138	Ce=140				
9								
10			Er=178	La=180	Ta=182	W=184		Os=195, Ir=197 Pt=198, Au=199
11		Hg=200	Tl=204	Pb=207	Bi=208			
12				Th=231		U=240		

William Ramsay & Rayleigh (1894)

In 1894 Lord Rayleigh gave a lecture about a discrepancy between the density of nitrogen chemically synthesised and isolated nitrogen from the known elements in the air. William Ramsay followed this up, what he found was an unknown heavy component of air. He observed that the gas had no obvious reactivity with other elements and named the gas argon. He also noticed that the gas had no place in Mendeleev's table of elements.

In 1898 the other noble gases; helium, neon, krypton, and xenon were discovered and recognised to follow a pattern, they were identified as a new group on the periodic table (group 8 or 0).

Henry Moseley (1914)

Henry Moseley proposed the concept of the atomic number (the number of protons in the nucleus or electrons orbiting the nucleus of a neutral atom) in 1914. Moseley also clarified several discrepancies in Mendeleev's table. Moseley determined the atomic numbers for each of the elements, and changed the periodic law to state that: - the atomic number of an element is the factor which determines chemical properties.

Predictions made after the Periodic Table's construction

As mentioned above (see Mendeleev's Entry), Mendeleev predicted the existence of several elements before they were discovered. The most famous of which was the prediction of eka-silicon or as it is now known as germanium. Mendeleev also managed to predict the properties of this element, which have been shown to be extremely accurate for his time (see table 1.4).

Table 1.4 *Comparison of Mendeleev's predictions and the properties of Germanium.*

Predicted properties of eka-silicon (X)		Properties of Germanium (Ge)
Colour	grey	Greyish-white
Atomic weight	72	72.59
Density (g/cm ³)	5.5	5.323
Formula for oxide	XO ₂	GeO ₂
Formula for chloride	XCl ₄	GeCl ₄
Melting point	Very High	937.4°C
Specific heat capacity	0.31 J K ⁻¹ g ⁻¹	0.32 J K ⁻¹ g ⁻¹

Part B

Trends across periods and down groups

Introduction

The periodic table in its modern form is a resource for the properties of the elements but also serves as a predictive tool. Across periods and down groups a range of physical and chemical properties emerge which show distinct trends. For science to have an understanding of these trends, it is important to explain these trends in terms of atomic structure.

First ionisation energy

Reference: 1:2:3:4:5:7

In order to remove an electron from an atom, energy is required. Ionisation energy is the amount of energy to remove electrons from atoms. The first ionisation energy is the energy required to remove one electron from the neutral atom.

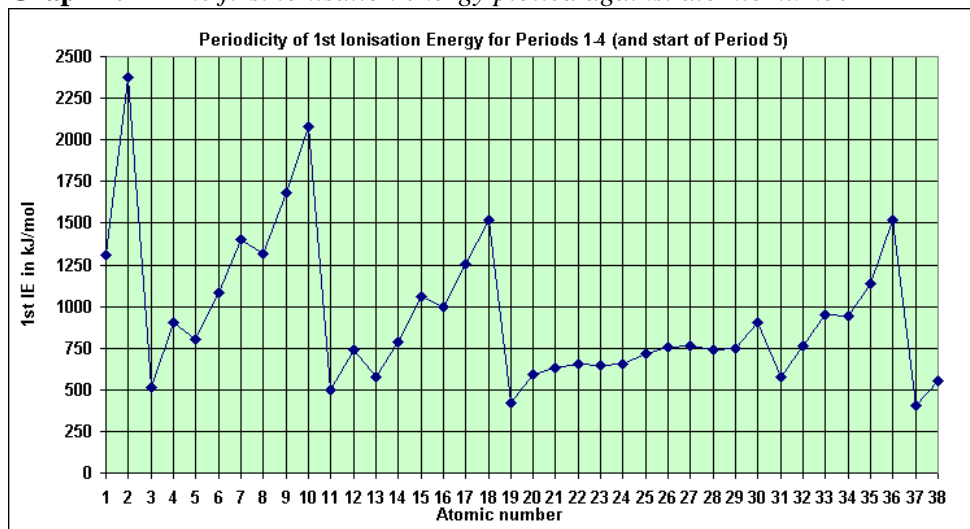
Despite some irregularities, ionisation energy increases across the periods of the periodic table. This is because of the valence electrons being closer to the nucleus for the elements on the right hand side of a period. Also the number of protons in the nucleus is increased, so there is greater force of attraction holding the outer electrons in place, thus more energy is required to remove the first electron.

Ionisation energy decreases down groups, because the valence electrons are further away from the nucleus because of the presence of additional shells of electrons. Also the inner shells shield the valence electrons from the positive charged nucleus. Therefore the force of attraction between the nucleus and the outer shell is reduced, thus less energy is required. The decrease in ionisation energy moving down groups renders these elements more reactive. Take the Group 1 alkali metals for example, Lithium reacts fairly gradually with water, but moving further down the group we have caesium which is so reactive it must be kept sealed from air.

The difference in ionisation energies suggests why metals always form cations and non-metals always form anions in ionic compounds. It also provides strong circumstantial evidence for periodic law.

Note that in graph 1.1 it is the noble gases which form the peaks of the graph and that the group 1 elements form the minima troughs. Group 1 elements have one more electron than the noble gases and therefore readily lose that electron to obtain the noble gas electron configuration. Noble gases have stable electron configurations and thus resist ionisation, which would make them unstable.

Graph 1.1 – The first ionisation energy plotted against atomic number



Reference: 1:3:4:5:7

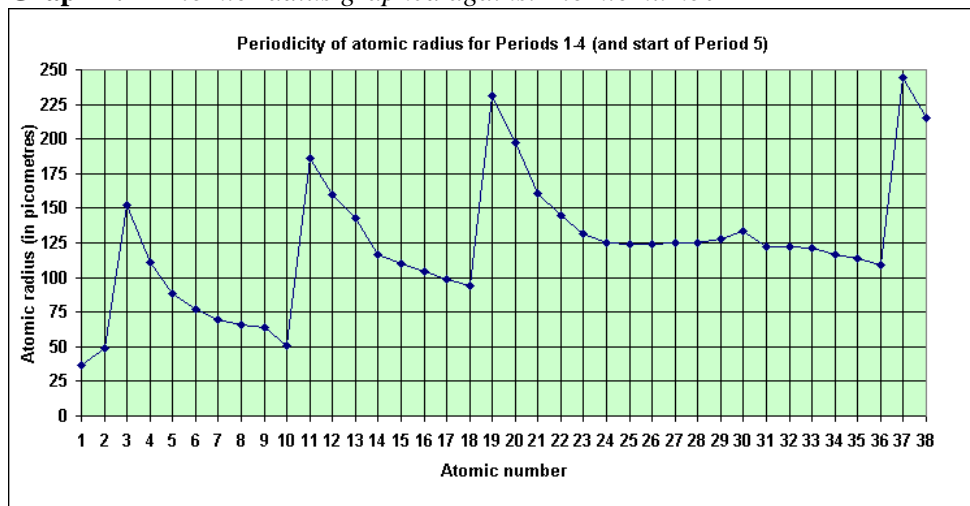
Atomic radius

Atomic radius is defined to be the distance between the nucleus of the atom and the outer most electron orbital, when the atom is in its neutral state. In metals atomic radius is half the distance between the nuclei in two adjacent atoms. For the elements which exist as diatomic molecules (eg. Chlorine) the atomic radius is half the distance between the nuclei of the two atoms within the molecule. The atomic radius decreases across periods as the nuclear charge increases. The greater the nuclear charge the greater the force of attraction between the electrons and the nucleus, and the radius decreases.

The atomic radius increases down groups since the number of shells of electrons increases with each period. As new shells are added the effect of the nuclear charge (force of attraction) is overcome, and so the radius increases.

In the graph below (graph 1.2) the peaks represent the alkali metals (group 1) and the troughs represent the halogens (group 7). This gives strong evidence to support that the alkali metals have an extra electron orbital and for this reason they tend to lose that electron to obtain noble gas configuration and thus gives evidence for periodic law.

Graph 1.2 - Atomic radius graphed against Atomic number



Reference: 1:2:3:4:5:7:

Electronegativity

The ability of an element to attract the electrons to itself in a chemical bond is called electronegativity. This trend goes hand in hand with *ionisation energy* and *electron affinity* (The ability of an atom to accept

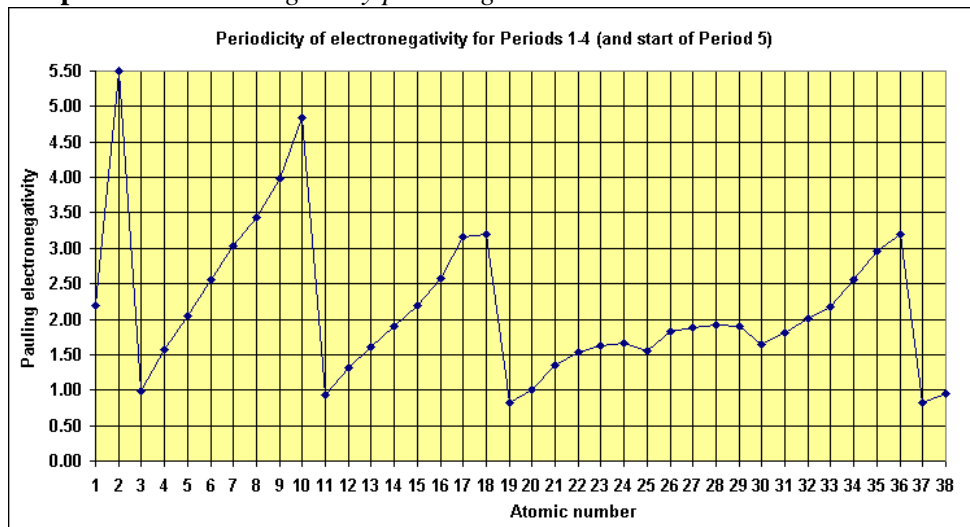
electrons), and plays a major roll in determining a compounds polarity. In practice electronegativity increases with moving left to right across periods, but decreases with moving down groups.

As mentioned above, this is largely determined by the ionisation energy and electron affinity of the element. Fluorine is a perfect example; this element has a high ionisation energy making it less tentative to loose electrons, and has a negative electron affinity value, which makes it tentative to receive electrons. Thus Fluorine has a high electronegativity value.

Caesium and Francium on the other end of the periodic table exhibit the lowest electronegativity values. This is because they are at the far left hand side of the table and at the very bottom of the group 1 metals.

The type of bond formed during a chemical reaction is largely determined by electronegativity. Atoms with equal electronegatives will tend to share electrons and hence form a covalent bond. If the electronegativity of one atom is greater or smaller then the other, the compound will be polar (ends slightly more negative/positive). Also if the electronegative difference is greater then 1.5/2.0 (slight discrepancy between values) then the compound will be ionic, and the elements will form ions.

Graph 1.3 – *Electronegativity potted against Atomic Number*



Note that in graph 1.3, the halogens have the highest electronegatives (i.e. the crests of the graph), and the group 1 alkali metals have the lowest electronegatives (i.e. the troughs). This explains why metals especially the group 1 metals, have a tendency to loose electrons, while the non-metals tend to gain electrons when participating in chemical reactions.

Reference: 1:3:4:5:7

Valency

Valency is the combining power of an atom or the number of electrons it gives or receives. Valency in its most common form of an element is the amount of electrons it needs to receive or loose in order to obtain octet configuration. Thus the group of an element determines its valency.

Group 0 (or 8) the 'noble gases' each have a full outer orbital, making them extremely stable and un-reactive, thus there valency is zero.

Group 1 the alkali metals each have one electron in there outer orbital and since it is easier for an atom to loose one then gain seven electrons, the valency is one.

The valencies of the transition metals are not so easy to predict, just from there positions alone. These metals mostly form 2+ cations, but can form cations with other charges, depending on the element the metal is bonded with. Examples include Cu^{2+} (copper II ion), Fe^{3+} (iron III ion) and Zn^{2+} (Zinc ion) which only form a 2+ charge.

The non-metals or less generally speaking; groups' five to seven tend to gain electrons and therefore they form anions (negative ions). The formula for there valencies is given by 8 minus the number of the group, for example group 5, has a valency of three.

Reference:3:4:9

Reactivity

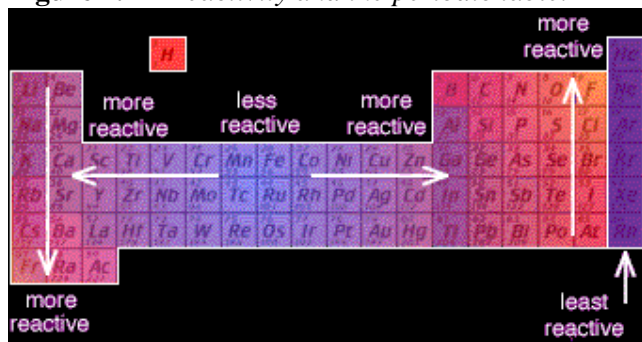
Reactivity refers to the rate at which a substance undergoes a chemical reaction. Reactivity is largely influenced by valency and electron affinity, since reactions take place in order to allow a substance to

enter a stable state. As such, the position of an element on the periodic table is vital in determining its reactivity.

For metals reactivity increases when moving down groups, and decreases when moving left to right across periods towards the centre of the periodic table, or more specifically the elements iron and cobalt. Metal reactivity also increases slightly when moving left from iron towards aluminium or group five.

For non-metals on the other hand, reactivity increases when moving vertically up groups and left to right across periods. The only exception is group five, the 'noble gases', which due to their stable nature, are extremely un-reactive.

Figure 1.1 – *Reactivity and the periodic table.*



- Figure 1.1 shows the reactivity of the elements of the periodic table.

Bibliography

Footnote Reference

Books:

1. C. M. Roebuck. Excel Preliminary Chemistry. Pascal Press, Glebe NSW. (2004).
2. C. M. Roebuck. Excel HSC Chemistry. Pascal Press, Glebe NSW. (2004).
3. Geoffrey Thickett. Pathways to Chemistry. Macmillan Education Australia, Melbourne. (1997).
4. Raymond Chang. Chemistry –Fourth Edition. McGraw Hill Book Company, Inc, USA. (1991).
5. Roland Smith. Preliminary Conquering Chemistry – Third Edition. McGraw Hill Book Company, Australia, Roseville NSW (2000).

Internet Sites:

6. **History of the Periodic Table of the Elements**. Available from
<www.auetute.com.au/ptistor.html> [accessed 17th April, 2006]
7. **Data for graphs**. <<http://www.wpbschoolhouse.btinternet.co.uk/page07/ASA2eledata38.htm>>
[accessed 19th April, 2006]
8. **Wikipedia**. Available from<http://en.wikipedia.org/wiki/Main_Page> [accessed 1st May, 2006]-
Use Search field for each trend.
9. **Reactivity and the periodic Table**. Available from:
<<http://www.nelsonthornes.com/secondary/science/scinet/scinet/reaction/react/periodic.htm>>
[accessed 2nd may, 2006]