Quadrant II – Transcript and Related Materials

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Module Name: Trends in Periodic Table for Ionisation Energy

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Notes:

Trends in Periodic Table for Ionisation Energy:

If a small amount of energy is supplied to an atom, then an electron may be promoted to a higher energy level, but if the amount of energy supplied is sufficiently large the electron may be completely removed. The energy required to remove the most loosely bound electron from an isolated gaseous atom is called the ionization energy. Ionization energies are determined from spectra and are measured in kJ $mo1^{-1}$. It is possible to remove more than one electron from most atoms. The first ionization energy is the energy required to remove the first electron and convert M to M^+; the second ionization energy is the energy required to remove the second election arid convert M^+ to M^{2+}; the third ionization energy converts M^{2+} to M^{3+}, and so on. The factors that influence the ionization energy are:

1. The size of the atom.
2. The charge on the nucleus.
3. How effectively the inner electron shells screen the nuclear charge.
4. The type of electron involved *(s. p, d* or/).

These factors are usually interrelated. In a small atom the electrons are tightly held. Whilst in a larger atom the electrons are less strongly held. Thus the ionization energy decreases as the size of the atoms increases.

Table 6.2 Ionization energies for Group I and II elements ($kJ\,mol^{-1}$)

	1st	2nd		1st	2nd	3rd
Li	520	7296	Be	899	1757	14847
Na	496	4563	Mg	737	1450	7731
K	419	3069	Ca	590	1145	4910
Rb	403	2650	Sr	549	1064	4207
Cs	376	2420	Ba	503	965	
Fr			Ra	509	979	3281*

*Estimated value.

Source: J. D. Lee, Concise Inorganic Chemistry

This trend is shown, for example, by Group I and Group II elements (See Table 6.2), and also by the other main groups. Comparison of the first and second ionization energies for the Group I elements shows that removal of a second electron involves a great deal more energy, between 7 and 14 times more than the first ionization energy. Because the second ionization energy is so high, a second electron is not removed. The large difference between the first and second ionization energies is related to the structure of the Group I atoms. These atoms have just one electron in their outer shell. Whilst it is relatively easy to remove the single outer electron, it requires much more energy to remove a second electron, since this involves breaking into a filled shell of electrons. The ionization energies for the Group II elements show that the first ionization energy is almost double the value for the corresponding Group I element. This is because the increased nuclear charge results in a smaller size for the Group II element. Once the first electron has been removed, the ratio of charges on the nucleus to the number' of orbital electrons (the effective nuclear charge) is increased, and this reduces the size. For example, Mg^+ is smaller than the Mg atom. Thus the remaining electrons in Mg^+ are even more tightly held, and consequently the second ionization energy is greater than the first. Removal of a third electron from a Group II element is very much harder for two reasons:

1. The effective nuclear charge has increased, and hence the remaining electrons are more tightly held.
2. Removing another electron would involve breaking a completed shell of electrons.

The ionization energy also depends on the type of electron which is removed *s, p, d* and *f* electrons have orbitals with different shapes. An *s* electron penetrates nearer to the nucleus, and is therefore more tightly held than a *p* electron. For similar reasons a *p* electron is more tightly held than a *d* electron, and a *d* electron is more tightly held than an *f* electron. Other factors being equal, the ionization energies are in the order $s > p > d > f$. Thus the increase in ionization energy is not quite smooth on moving from left to right in the periodic table.

Table 6.3 Comparison of some first ionization energies (kJ mol^{-1})

Li 520	Be 899	B 801	C 1086	N 1403	O 1410	F 1681	Ne 2080
Na 496	Mg 737	Al 577	Si 786	P 1012	S 999	Cl 1255	Ar 1521

Group / Period	I	II											III	IV	V	VI	VII	0
1	H 1311																	He 2372
2	Li 520	Be 899											B 801	C 1086	N 1403	O 1410	F 1681	Ne 2080
3	Na 496	Mg 737											Al 577	Si 786	P 1012	S 999	Cl 1255	Ar 1521
4	K 419	Ca 590	Sc 631	Ti 656	V 650	Cr 652	Mn 717	Fe 762	Co 758	Ni 736	Cu 745	Zn 906	Ga 579	Ge 760	As 947	Se 941	Br 1142	Kr 1351
5	Rb 403	Sr 549	Y 616	Zr 674	Nb 664	Mo 685	Tc 703	Ru 711	Rh 720	Pd 804	Ag 731	Cd 876	In 558	Sn 708	Sb 834	Te 869	I 1191	Xe 1170
6	Cs 376	Ba 503	La 541	Hf 760	Ta 760	W 770	Re 759	Os 840	Ir 900	Pt 870	Au 889	Hg 1007	Tl 589	Pb 715	Bi 703	Po 813	At 912	Rn 1037
7	Fr	Ra	Ac															

FIRST IONIZATION ENERGIES OF THE ELEMENTS
(Numerical values are given in kJ mol^{-1}.)
(Large circles indicate high values and small circles low values.)
After Sanderson, R.T., *Chemical Periodicity*, Reinhold, New York.

Source: J. D. Lee, Concise Inorganic Chemistry

For example, the first ionization energy for a Group III element (where a *p* electron is being removed) is actually less than that for the adjacent Group II element (where an *s* electron is being removed). In general, the ionization energy decreases on descending a group and increases on crossing a period. Removal of successive electrons becomes more difficult and first ionization energy < second ionization energy < third ionization energy. There are a number of deviations from these generalizations.

The variation in the first ionization energies of the elements are shown in Figure 6.1. The graph shows three features:

1. The noble gases He, Ne, Ar, Kr, Xe and Rn have the highest ionization energies in their respective periods.
2. The Group I metals Li, Na, K and Rb have the lowest ionization energies in their respective periods.
3. There is a general upward trend in ionization energy within a horizontal period, for example from Li to Ne or from Na to Ar.

The values for Ne and Ar are the highest in their periods because a great deal of energy is required to remove an electron from a stable filled shell of electrons.

The graph does not increase smoothly. The values for Be and Mg are high, and this is attributed to the stability of a filled *s* level. The values for N and P are also high, and this indicates that a half filled *p* level is also particularly stable. The values for B and Al are lower because removal of one electron leaves a stable filled *s* shell, and similarly with O and S a stable half filled *p* shell is left.

In general, the first ionization energy decreases in a regular way on descending the main groups. A departure from this trend occurs in Group III, where the expected decrease occurs between B and Al, but the values for the remaining elements Ga, In and Tl do not continue the trend, and are irregular. The reason for the change at Ga is that it is preceded by ten elements of the first transition series (where the 3d shell is being filled). This makes Ga smaller than it would otherwise be. A similar effect is observed with the second and third transition series, and the presence of the three transition series not only has a marked effect on the values for Ga, In and Tl, but the effect still shows in Groups IV and V.

The ionization energies of the transition elements are slightly irregular, but the third row elements starting at Hf have lower values than would be expected due to the interpolation of the 14 lanthanide elements between La and Hf.

Table 6.5 Ionization energies for Group III elements (kJ mol^{-1})

	1st	2nd	3rd
B	801	2427	3659
Al	577	1816	2744
Ga	579	1979	2962
In	558	1820	2704
Tl	589	1971	2877