Chapter 8: Electron Configuration + Periodicity

Electron Configuration of Atoms

– basis for explaining the behavior of elements in periodic table.

Nomenclature:

__	atomic orbital
↑	unpaired electron
↑↓	spin paired electrons

Note: The Pauli Exclusion Principle says that ↑↑ is impossible.

First Period (1st horizontal row) of Periodic Table: $n=1$

	1s	Notation
$_1H$	↑	$1s^1$
$_2He$	↑↓	$1s^2$

Recall: for a neutral atom, the atomic number (# protons) = # electrons

He is a very stable element – noble gas. $n=1$ energy level filled

Second Period (2nd horizontal row) of Periodic Table: $n=2$

	1s	2s	2p			Shorthand Notation
$_3Li$	↑↓	↑	__	__	__	$1s^2\ 2s^1$
$_4Be$	↑↓	↑↓	__	__	__	$1s^2\ 2s^2$
$_5B$	↑↓	↑↓	↑	__	__	$1s^2\ 2s^2\ 2p^1$
$_6C$	↑↓	↑↓	↑	↑	__	$1s^2\ 2s^2\ 2p^2$
$_7N$	↑↓	↑↓	↑	↑	↑	$1s^2\ 2s^2\ 2p^3$ or $1s^2\ 2s^2\ 2p_x^1\ 2p_y^1\ 2p_z^1$
$_8O$	↑↓	↑↓	↑↓	↑	↑	$1s^2\ 2s^2\ 2p^4$
$_9F$	↑↓	↑↓	↑↓	↑↓	↑	$1s^2\ 2s^2\ 2p^5$
$_{10}Ne$	↑↓	↑↓	↑↓	↑↓	↑↓	$1s^2\ 2s^2\ 2p^6$ very stable $n=1$, $n=2$ energy levels filled

$\underbrace{\quad 2s \quad 2p \quad}_{\#e\text{'s} = 2n^2 = 8}$

We used Hund's Rule: electrons must occupy all the orbitals of a given sublevel singly before pairing begins.

diamagnetic: refers to an atom in which the electrons are all paired up
examples: $_2He$, $_4Be$, $_{10}Ne$

paramagnetic: refers to an atom in which there are unpaired electrons.
examples: $_5B$, $_9F$ (1 unpaired e^-) $_7N$ (3 unpaired e's) – exhibits paramagnetism to greatest extent.
C, O (2 unpaired e⁻'s)

Third Period (3rd horizontal row) of Periodic Table : n=3

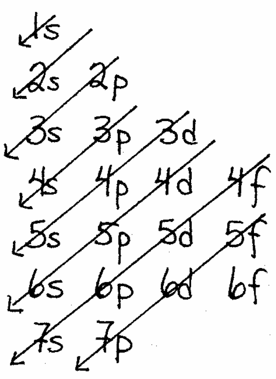
	10e⁻	3s	3p			Shorthand Notation
$_{11}$Na	[Ne]	↑				[Ne] $3s^1$ or $1s^2 2s^2 2p^6 3s^1$
$_{12}$Mg	[Ne]	↑↓				[Ne] $3s^2$
$_{13}$Al	[Ne]	↑↓	↑			[Ne] $3s^2 3p^1$
$_{14}$Si	[Ne]	↑↓	↑	↑		[Ne] $3s^2 3p^2$
$_{15}$P	[Ne]	↑↓	↑	↑	↑	[Ne] $3s^2 3p^3$
$_{16}$S	[Ne]	↑↓	↑↓	↑	↑	[Ne] $3s^2 3p^4$
$_{17}$Cl	[Ne]	↑↓	↑↓	↑↓	↑	[Ne] $3s^2 3p^5$
$_{18}$Ar	[Ne]	↑↓	↑↓	↑↓	↑↓	[Ne] $3s^2 3p^6$ very stable

Fourth Period (4th horizontal Row) of Periodic Table: n=4.

	18e⁻	3d					4s	4p		
$_{19}$K	[Ar]						↑			
$_{20}$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	↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

Exception to Filling Rule

What we saw was that the energy levels filled up differently than what we expected. One way to remember the general order is



Note: there are still exceptions to the filling order eg Cu & Cr. These two are the only ones you are responsible for.

Another way is to use the periodic table.

The periodic table has A groups and B groups. (vertical columns)

(Note: periodic tables can be different - use periodic table on sheet + in textbook

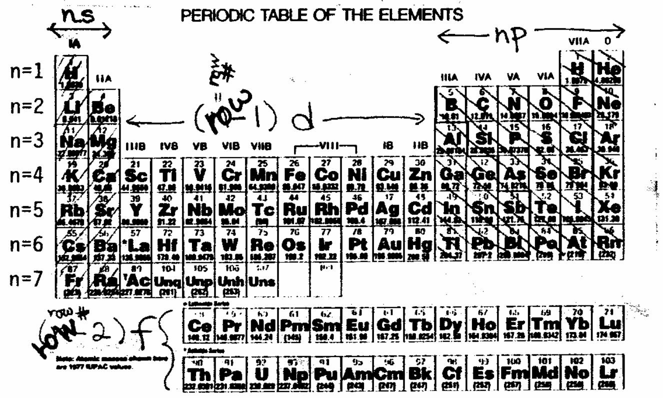
Periods are horizontal rows: 1st period corresponds to n=1 energy level etc.

A Group: s and p orbitals are being filled.

B Group: ns orbital is filled and (n-1) d sublevel is being filled.

Exercise: Follow the elements in periodic table in order of increasing atomic number and state what orbitals are being filled. (~~Table 5-5~~) Figure 5-31

Example: In the fourth period (n=4), we first put electrons in 4s orbital, then electrons are placed in n-1 = 4-1 = 3d sublevel in B groups. Finally electrons are put into 4p energy level.



ELECTRONIC CONFIGURATION OF THE ELEMENTS

There are two elements in the first period (row) of the periodic table, hydrogen and helium. The electrons in hydrogen and helium atoms occupy the 1s orbital.

	1s
$_1$H	—
$_2$He	—

There are eight elements (numbers 3-10) in the second period of the periodic table. The outermost electrons in these elements occupy the 2s and 2p orbitals; that is, they all have n = 2 where n is the outermost occupied energy level.

	1s	2s	2p	Outermost Configuration
$_3$Li	—	—	— — —	
$_4$Be	—	—	— — —	
$_5$B	—	—	— — —	
$_6$C	—	—	— — —	
$_7$N	—	—	— — —	
$_8$O	—	—	— — —	
$_9$F	—	—	— — —	
$_{10}$Ne	—	—	— — —	

There are also eight elements (numbers 11-18) in the third period. The outermost electrons in these elements occupy the 3s and 3p orbitals (n = 3).

		3s	3p
$_{11}$Na	[Ne]	—	— — —
$_{12}$Mg	[Ne]	—	— — —
$_{13}$Al	[Ne]	—	— — —
$_{14}$Si	[Ne]	—	— — —
$_{15}$P	[Ne]	—	— — —
$_{16}$S	[Ne]	—	— — —
$_{17}$Cl	[Ne]	—	— — —
$_{18}$Ar	[Ne]	—	— — —

The symbol for a noble gas indicates a number of electrons corresponding to the atomic number of the noble gas, in filled sets of orbitals; [He] corresponds to 2 e^-, and [Ar] corresponds to 18 e^- in filled sets of orbitals.

There are eighteen elements (numbers 19-36) in the fourth period of the periodic table. As we shall see, the electrons added in these elements occupy the 4s, 3d, and 4p orbitals. (N = 4)

		3d	4s	4p
$_{19}$K	[Ar]	— — — — —	—	— — —
$_{20}$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		— — — — —	—	— — —

[Ar] is not repeated 17 time. It is understood to be there

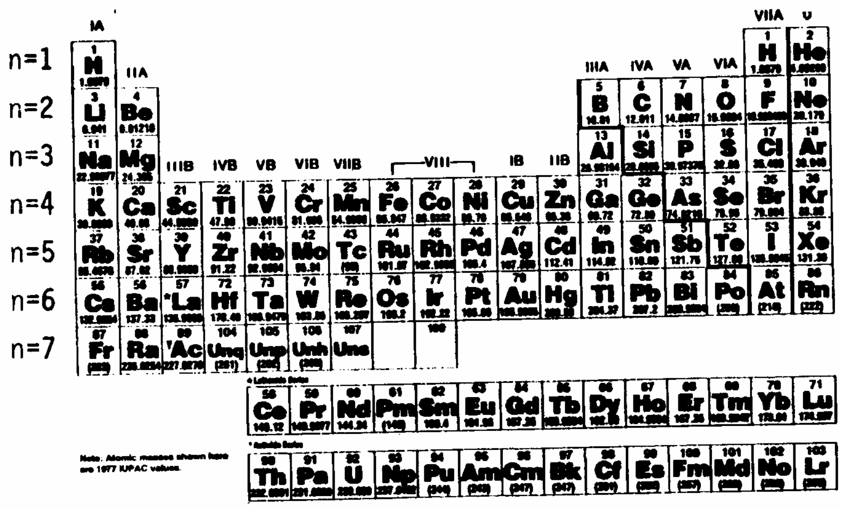
The pattern for filling atomic orbitals has been evolved.

Summary on Order of Filling Atomic Orbitals

There are several mnemonic devices (including this one) to assist one in remembering the order in which electrons are "added" to atoms.

1s
2s 2p
3s 3p 3d
4s 4p 4d 4f
5s 5p 5d 5f
6s 6p 6d
7s 7p

PERIODIC TABLE OF THE ELEMENTS



	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VIII	VIII	VIII	IB	IIB	IIIA	IVA	VA	VIA	VIIA	0
n=1	1 H																1 H	2 He
n=2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
n=3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
n=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
n=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
n=6	55 Cs	56 Ba	57 *La	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
n=7	87 Fr	88 Ra	89 †Ac	104 Unq	105 Unp	106 Unh	107 Uns		109									

* Lanthanide Series

58	59	60	61	62	63	64	65	66	67	68	69	70	71
Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er	Tm	Yb	Lu

† Actinide Series

90	91	92	93	94	95	96	97	98	99	100	101	102	103
Th	Pa	U	Np	Pu	Am	Cm	Bk	Cf	Es	Fm	Md	No	Lr

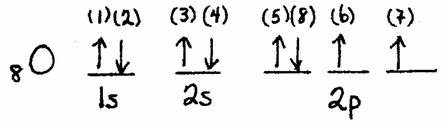
Note: Atomic masses shown here are 1977 IUPAC values.

Example: In the sixth period (n = 6), the 6s orbital is filled, the the energy sublevels are filled in the following order: 4f, 5d then 6p.

Relation Between Quantum Numbers and Electronic Configuration

Recall: every election in a atom has its own unique set of 4 quantum numbers.

Example: Write a complete set of 4 quantum numbers for each election in an oxygen atom, $_8O$. There are 8 electrons in $_8O$; there should be 8 sets of 4 quantum numbers, one for each election.



Election	n		ℓ		m_ℓ	m_s
1	1	($\underset{\uparrow n}{1}\underset{\nwarrow \ell}{s}$ orbital)	0	(since is s orbital)	0	$+\frac{1}{2}$
2	1		0		0	$-\frac{1}{2}$
3	2		0	(since is s orbital)	0	$+\frac{1}{2}$
4	2		0		0	$-\frac{1}{2}$
5	2		1	(since is p orbital)	-1	$+\frac{1}{2}$
6	2		1		0	$+\frac{1}{2}$
7	2		1		+1	$+\frac{1}{2}$
8	2		1		-1	$-\frac{1}{2}$

Notes on bookkeeping: ↑ ≡ spin is $+\frac{1}{2}$

↓ ≡ spin is $-\frac{1}{2}$

the 3 p orbitals are labelled $m_\ell =$ $\underline{-1}$ $\underline{0}$ $\underline{+1}$

Example: What is an appropriate set of 4 quantum numbers for the last election to go into an element of:

		n	ℓ	m_ℓ	m_s			n	ℓ	m_ℓ	m_s
(a)	Na	3	0(s)	0	$\pm\frac{1}{2}$	(c)	$_{82}Pb$	6	1(p)	-1,0,+1	$\pm\frac{1}{2}$
(b)	Ru	4	2(d)	-2,-1,0 +1 +2	$\pm\frac{1}{2}$	(d)	No	5	3(f)	-3,-2,-1,0 1,2,3	$\pm\frac{1}{2}$

Preview of Periodic Properties

The elements of a certain Group tend to behave similarly in chemical reactions because it is the OUTSIDE electrons (the elections that are farthest from the nucleus) that are involved in chemical reactions These are the outermost s and p elections – called VALENCE elections. The elements in a certain group all have the same number of valence electrons.

Group IA (H, Li, Na, K····) have 1 electron in outermost s orbital: ns^1
Group II A (Be, Mg, Ca ····) have 2 elections in outermost s orbital: ns^2
Group VIA (O, S, Se ····) have 6 electrons in outermost s+p orbitals: ns^2np^4

Summerizing Example:

(a) The not-yet-discovered element with atomic number 115 falls into which periodic group?
(b) Its chemical behavior would be most similar to which elements?
(c) Its outer elections would have what shorthand notation? How many valence e's?
(d) Its election configuration is what? (use periodic table)
(e) An acceptable set of 4 quantum numbers for the last electron to be put into this element is ___.

ANS:

(a) VA
(b) N, P, As – other VA elements
(c) $7s^2\,7p^3$; 5 valence elections
(d) $\underset{86e^-}{[Rn]}\ 5f^{14}\,6d^{10}\,7s^2\,7p^3$

(e) $n = 7$, $\ell = 1$, m_ℓ could be -1 or 0 or $+1$, m_s could be $+\frac{1}{2}$ or $-\frac{1}{2}$

Chemical Periodicity

We just finished discussing the structure of atoms and introduced its relationship to chemical periodicity: the variation in the properties of elements with their positions in the periodic table.

electron as wave-like ⇒ Schrodinger Equation ⇒ quantum numbers: n, ℓ, m_ℓ, m_s ⇒ electron configuration $1s^2 2s^2 2p^6 \ldots$ ⇒ predicted properties (periodicity) according to periodic table

The history of the periodic table is given in the text .. Mendeleev and Meyer (1869) independently arranged the known elements into a pattern based on their properties, thus setting the stage for the ....

Periodic Law : The properties of the elements are periodic functions of their atomic numbers (the numbers of protons in the nucleus).

Terms for the Periodic Table:

group or family : vertical columns
the elements have similar chemical and physical properties.

period : horizontal rows
the elements have properties that change progressively across the table.

Group A: "representative" elements
show distinct and fairly regular variations in properties with the atomic number
their last electron is added to an s or p orbital

Group B: show less dramatic changes in properties with increasing atomic number
their last electron is added to an (n-1) d orbital which lies inside the outer ns and np electrons.

The representative elements:

Group IA (except H)	alkali metals	ns^1
Group IIA	alkaline earth metals	ns^2
Group VIIA	halogens (salt-formers)	ns^2np^5
Group 0	noble (rare) gases	ns^2np^6 (except He)

d-transition elements: in B group (except IIB)
electrons are being added to d-orbitals
(Note: IIB is excluded since their last e^- is added to s orbital. Elements are Zn, Cd, Hg, but their properties are similar to d-transition elements.

1st transition series: $_{21}Sc \rightarrow _{29}Cu$
2nd transition series: $_{39}Y \rightarrow _{47}Ag$
3rd transition series: $_{57}La$, $_{72}Hf \rightarrow _{79}Au$
4th transition series: $_{89}Ac$

Inner transition elements: located between IIIB and IVB
electrons are being added to f orbitals
1st inner transition series: lanthanides $_{58}Ce \rightarrow _{71}Lu$ (lutetium)
2nd inner transition series: actinides $_{90}Th \rightarrow _{103}Lr$ (lawrencium)

Lewis Dot Representations of Representative Elements

- only the elections shown in the outer s & p orbitals are shown as dots
  These are called VALENCE electrons.
- paired and unpaired electrons are shown.

TABLE ~~5-3~~ 7-1 Lewis Electron Dot Formulas for Representative Elements

GROUP	IA	IIA	IIIA	IVA	VA	VIA	VIIA	0
Number of Electrons in Outer Shell	1	2	3	4	5	6	7	8 (except He)
Row 1	$\mathrm{H}\cdot$							$\mathrm{He}:$
Row 2	$\mathrm{Li}\cdot$	$\mathrm{Be}:$	$\ddot{\mathrm{B}}\cdot$	$\underset{\cdot}{\ddot{\mathrm{C}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{N}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{O}}}:$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{F}}}:$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ne}}}:$
Row 3	$\mathrm{Na}\cdot$	$\mathrm{Mg}:$	$\ddot{\mathrm{Al}}\cdot$	$\underset{\cdot}{\ddot{\mathrm{Si}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{P}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{S}}}:$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Ar}}}:$
Row 4	$\mathrm{K}\cdot$	$\mathrm{Ca}:$	$\ddot{\mathrm{Ga}}\cdot$	$\underset{\cdot}{\ddot{\mathrm{Ge}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{As}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{Se}}}:$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{Br}}}:$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Kr}}}:$
Row 5	$\mathrm{Rb}\cdot$	$\mathrm{Sr}:$	$\ddot{\mathrm{In}}\cdot$	$\underset{\cdot}{\ddot{\mathrm{Sn}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{Sb}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{Te}}}:$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{I}}}:$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Xe}}}:$
Row 6	$\mathrm{Cs}\cdot$	$\mathrm{Ba}:$	$\ddot{\mathrm{Tl}}\cdot$	$\underset{\cdot}{\ddot{\mathrm{Pb}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{Bi}}}\cdot$	$\cdot\underset{\cdot}{\ddot{\mathrm{Po}}}:$	$\cdot\underset{\cdot\cdot}{\ddot{\mathrm{At}}}:$	$:\underset{\cdot\cdot}{\ddot{\mathrm{Rn}}}:$
Row 7	$\mathrm{Fr}\cdot$	$\mathrm{Ra}:$						

Memorize or
deduce from
periodic table.

Periodic Properties for Group A Elements

The physical and chemical properties of elements recur periodically.
There are 5 properties we will consider in depth

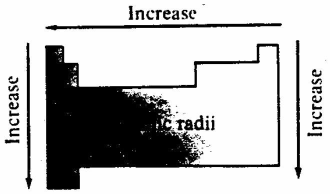
(1) atomic radii
(2) ionization energy
(3) electron affinity
(4) ionic radii
(5) electronegativity

(1) Atomic Radii

values have been determined indirectly (units : angstroms $1 Å = 10^{-10} m$
nanometers $1 \times 10^{-9} m$

as we move from left to right within a particular energy level (eg $n = 2$) the atomic number increases. Therefore for a neutral atom, protons are added to the nucleus increasing the nuclear charge; electrons are added to the electron cloud in same major energy level. There is greater and greater attraction; the electron cloud is pulled closer and closer into the nucleus.
Result : the atomic radii decreases. Eg Li > Be > B > C > N > O > F > Ne

as we move from top ($n=1$) to bottom ($n=7$), electrons are added to higher and higher energy levels which are correspondingly larger and larger. Result : the atomic radii increases. Eg Be < Mg < Ca < Sr < Ba



general trend in atomic radii of A group elements with position in periodic table.

Chemical reactions result from the interaction of the outermost electrons of the elements involved. Therefore it is important to know how tightly the electrons of an element are held and how easily the elements gain or lose electrons.

To give us this information, the energy involved in these processes is measured.

Convention : energy is released (exothermic) : energy has negative value
energy is absorbed (endothermic) : energy has positive value

(2) <u>First Ionization Energy</u>: minimum amount of energy required to remove the most loosely bound electron from an isolated gaseous atom to form an ion with a 1+ charge.

e.g. $K(g) \rightarrow K^+(g) + e^-$ I.E. = 419 kJ/mol of K(g) atoms

or $K(g) + 419\ kJ/mol \rightarrow K^+(g) + e^-$
(process requires energy – is endothermic)

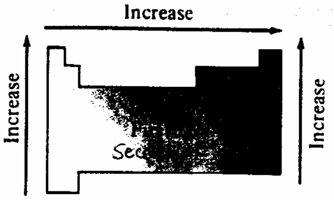
For $^{39}_{19}K^+$,
#p 19
#e 18 (it lost an e^-)
#n 20

For calcium:

First ionization energy, IE_1 $Ca(g) + 590\ kJ \rightarrow Ca^+(g) + e^-$
Second ionization energy, IE_2 $Ca^+(g) + 1145\ kJ \rightarrow Ca^{2+}(g) + e^-$

$IE_2 > IE_1$ since it is more difficult to remove an electron from a positively charged species than a neutral one

Again, IE measures how tightly the electrons are bound to the atom. As IE ↑, the difficulty in removing an electron ↑



General trend in IE_1 of A Group Elements with position in periodic table. Exceptions occur at IIIA and VIA. These can be explained by quantum mechanics (see pp 226)

∴ the outer He electron is more tightly bound than any other (has highest IE_1)
the outer Cs electron is the easiest to remove (has lowest IE_1)
(Note: Fr is ignored because it is rare and radioactive)

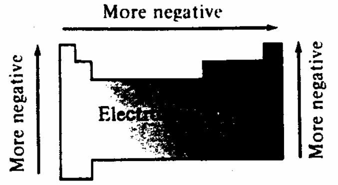
Knowledge of the relative values of IE_1 assist in predicting which elements form covalent bonds (by sharing electrons) and which form ionic bonds (electrons are transferred from one atom to another)

(3) <u>Electron Affinity</u> : the amount of energy involved in the process in which an electron is added to an isolated gaseous atom to form an ion with a 1- charge.

examples: $Be(g) + e^- + \overset{0}{\cancel{241}}\ kJ \rightarrow \underset{\text{anion}}{Be^-(g)}$ E.A. = $\overset{0}{+\cancel{241}}$ kJ/mol (~~endo~~thermic)

$Cl(g) + e^- \longrightarrow \underset{\text{anion}}{Cl^-(g)} + 348\ kJ$ E.A. = −348 kJ/mol (exothermic

for $^{35}_{17}Cl^-$ #p = 17
#e = 18 (charge of −1 means 1 extra electron)
#n = 18



General trend in electron affinities of A group elements with position in periodic table. There are many exceptions.

∴ it is easier to add an electron going from left to right or bottom to top.
∴ halogens gain electrons easily – have more negative electron affinities.

(4) <u>Ionic Radii</u>

Most parent atoms will react with other atoms in order to gain a noble gas configuration, by either gaining or losing electrons.

LEFT SIDE OF PERIODIC TABLE

Group	IA	IIA	IIIA
electronic config^n	ns^1	ns^2	ns^2np^1
electrons lost	1	2	3
final charge on ion	+1	+2	+3

Atoms in these groups will lose electrons to form cations and become positively charged.

In periods 2 and 3, ions from IA, IIA and IIIA are ISOELECTRONIC (have the same total number of electrons).
eg $_{11}Na^+$, $_{12}Mg^{2+}$ and $_{13}Al^{3+}$ all have <u>10 electrons</u> and are isoelectronic with the noble gas, $_{10}Ne$, neon.

In order of decreasing ionic radii : $Na^+ > Mg^{2+} > Al^{3+}$ since as more electrons are removed, the remaining electrons are held more tightly.
Also, $Na > Na^+$, $Mg > Mg^{2+}$

RIGHT SIDE OF PERIODIC TABLE

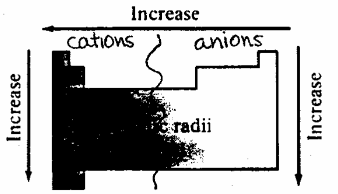
Group	VA	VIA	VIIA
electronic config^n	ns^2np^3	ns^2np^4	ns^2np^5
electrons gained	3	2	1
final charge on ion	-3	-2	-1

Atoms in these groups will gain electrons to have noble gas configuration - they form anions

Example: In period 2, ${}_7N^{3-}$, ${}_8O^{2-}$ and ${}_9F^-$ all have 10 electrons and are isoelectronic with ${}_{10}Ne$, ${}_{11}Na^+$, ${}_{12}Mg^{2+}$ and ${}_{13}Al^{3+}$.

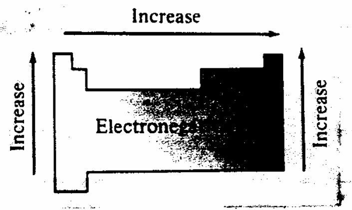
In order of decreasing ionic radii $N^{3-} > O^{2-} > F^-$. As electrons are added to a neutral atom, the electron cloud is expanded. Therefore $N^{3-} > N$

To generalize : eg $H^- > H > H^+$
(1) simple negatively charged ions (anions) are always larger than parent atom.
(2) simple positively charged ions (cations) are always smaller than parent atom.
(3) within an isoelectronic series, atomic radii decrease as atomic no. increases.



General trend in ionic radii of A Group elements with position in periodic table.

(5) Electronegativity: a measure of the relative tendency of an atom to attract electrons to itself when chemically combined with another atom. (These values appear on the periodic table in the exam)



General trend in electronegativity of A Groups with position in periodic table.

F highest EN (4.0)
O next highest EN (3.5)

Range
0.8 ⟶ 4.0
Fr F

Electronegativity is used to help assign oxidation numbers.

In a simple binary (2 element) compound, the more electronegative element "wants" electrons more and "gets" the negative oxidation number.

Examples:

no.	+1	-1	+1	-1	-3	+1
	Na	Cl	Na	H	N	H_3
N	1.0	3.0	1.0	2.1	3.0	2.1

Note: configuration of transition metal ions

Fe $[Ar]\ 3d^6\ 4s^2$ recall 4s has lower energy than 3d

Fe^{2+} $[Ar]\ 3d^6$ remove e^- from <u>outermost</u> orbitals first

Fe^{3+} $[Ar]\ 3d^5$

So, the order you put e^- into orbitals (order of increasing energy) is different from how you remove e^- from orbitals.