CHEM1011 LECTURE 11

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GROUND STATE ELECTRON CONFIGURATIONS

Group number

Row number	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 **H** $1s^1$																	2 **He** $1s^2$
	main group (s block)												main group (p block)					
2	3 **Li** $2s^1$	4 **Be** $2s^2$											5 **B** $2s^2 2p^1$	6 **C** $2s^2 2p^2$	7 **N** $2s^2 2p^3$	8 **O** $2s^2 2p^4$	9 **F** $2s^2 2p^5$	10 **Ne** $2s^2 2p^6$
3	11 **Na** $3s^1$	12 **Mg** $3s^2$	transition metals (d block)										13 **Al** $3s^2 3p^1$	14 **Si** $3s^2 3p^2$	15 **P** $3s^2 3p^3$	16 **S** $3s^2 3p^4$	17 **Cl** $3s^2 3p^5$	18 **Ar** $3s^2 3p^6$
4	19 **K** $4s^1$	20 **Ca** $4s^2$	21 **Sc** $4s^2 3d^1$	22 **Ti** $4s^2 3d^2$	23 **V** $4s^2 3d^3$	24 **Cr** $4s^1 3d^5$	25 **Mn** $4s^2 3d^5$	26 **Fe** $4s^2 3d^6$	27 **Co** $4s^2 3d^7$	28 **Ni** $4s^2 3d^8$	29 **Cu** $4s^1 3d^{10}$	30 **Zn** $4s^2 3d^{10}$	31 **Ga** $4s^2 4p^1$	32 **Ge** $4s^2 4p^2$	33 **As** $4s^2 4p^3$	34 **Se** $4s^2 4p^4$	35 **Br** $4s^2 4p^5$	36 **Kr** $4s^2 4p^6$
5	37 **Rb** $5s^1$	38 **Sr** $5s^2$	39 **Y** $5s^2 4d^1$	40 **Zr** $5s^2 4d^2$	41 **Nb** $5s^1 4d^4$	42 **Mo** $5s^1 4d^5$	43 **Tc** $5s^2 4d^5$	44 **Ru** $5s^1 4d^7$	45 **Rh** $5s^1 4d^8$	46 **Pd** $4d^{10}$	47 **Ag** $5s^1 4d^{10}$	48 **Cd** $5s^2 4d^{10}$	49 **In** $5s^2 5p^1$	50 **Sn** $5s^2 5p^2$	51 **Sb** $5s^2 5p^3$	52 **Te** $5s^2 5p^4$	53 **I** $5s^2 5p^5$	54 **Xe** $5s^2 5p^6$
6	55 **Cs** $6s^1$	56 **Ba** $6s^2$	57–71 *	72 **Hf** $6s^2 5d^2$	73 **Ta** $6s^2 5d^3$	74 **W** $6s^2 5d^4$	75 **Re** $6s^2 5d^5$	76 **Os** $6s^2 5d^6$	77 **Ir** $6s^2 5d^7$	78 **Pt** $6s^1 5d^9$	79 **Au** $6s^1 5d^{10}$	80 **Hg** $6s^2 5d^{10}$	81 **Tl** $6s^2 6p^1$	82 **Pb** $6s^2 6p^2$	83 **Bi** $6s^2 6p^3$	84 **Po** $6s^2 6p^4$	85 **At** $6s^2 6p^5$	86 **Rn** $6s^2 6p^6$
7	87 **Fr** $7s^1$	88 **Ra** $7s^2$	89–103 **	104 **Rf** $7s^2 6d^2$	105 **Db** $7s^2 6d^3$	106 **Sg** $7s^2 6d^4$	107 **Bh** $7s^2 6d^5$	108 **Hs** $7s^2 6d^6$	109 **Mt** $7s^2 6d^7$	110 **Ds** $7s^2 6d^8$	111 **Rg** $7s^1 6d^{10}$	112 **Cn** $7s^2 6d^{10}$	113 **Uut** $7s^2 7p^1$	114 **Uuq** $7s^2 7p^2$	115 **Uup** $7s^2 7p^3$	116 **Uuh** $7s^2 7p^4$	117 **Uus** $7s^2 7p^5$	118 **Uuo** $7s^2 7p^5$

lanthoids and actinoids (f block)

6	lanthanoid series *	57 **La** $6s^2 5d^1$	58 **Ce** $6s^2 4f^1 5d^1$	59 **Pr** $6s^2 4f^3$	60 **Nd** $6s^2 4f^4$	61 **Pm** $6s^2 4f^5$	62 **Sm** $6s^2 4f^6$	63 **Eu** $6s^2 4f^7$	64 **Gd** $6s^2 4f^7 5d^1$	65 **Tb** $6s^2 4f^9$	66 **Dy** $6s^2 4f^{10}$	67 **Ho** $6s^2 4f^{11}$	68 **Er** $6s^2 4f^{12}$	69 **Tm** $6s^2 4f^{13}$	70 **Yb** $6s^2 4f^{14}$	71 **Lu** $6s^2 4f^{14} 5d^1$
7	actinoid series **	89 **Ac** $7s^2 6d^1$	90 **Th** $7s^2 6d^2$	91 **Pa** $7s^2 5f^2 6d^1$	92 **U** $7s^2 5f^3 6d^1$	93 **Np** $7s^2 5f^4 6d^1$	94 **Pu** $7s^2 5f^6$	95 **Am** $7s^2 5f^7$	96 **Cm** $7s^2 5f^7 6d^1$	97 **Bk** $7s^2 5f^9$	98 **Cf** $7s^2 5f^{10}$	99 **Es** $7s^2 5f^{11}$	100 **Fm** $7s^2 5f^{12}$	101 **Md** $7s^2 5f^{13}$	102 **No** $7s^2 5f^{14}$	103 **Lr** $7s^2 5f^{14} 6d^1$

ANOMALIES

Electrons in the 3*d* orbitals are **more stable when exactly half-filled ($3d^5$) and completely filled ($3d^{10}$)**, so this will be more energetically favourable than having a filled 4s orbital.

Chromium and **copper** are the anomalies in the 3d sublevel (there are others in later periods of the periodic table).

23	24	25	26	27	28	29	30
V	**Cr**	**Mn**	**Fe**	**Co**	**Ni**	**Cu**	**Zn**
$4s^2 3d^3$	$4s^1 3d^5$	$4s^2 3d^5$	$4s^2 3d^6$	$4s^2 3d^7$	$4s^2 3d^8$	$4s^1 3d^{10}$	$4s^2 3d^{10}$

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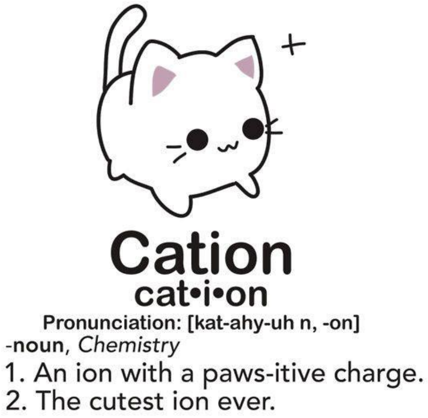
IONS

Anions

A neutral atom becomes an anion by *gaining* an electron – anions are **negatively** charged ions.



Cations

A neutral atom becomes a cation by *losing* an electron – cations are **positively** charged ions.

Often main group elements will lose or gain electrons to be become **isoelectronic** (same electron configuration) with the nearest noble gas.

ION CONFIGURATIONS

Anions

To determine the electronic configuration of an anion simply **add electrons following the *aufbau* principle** (as before when increasing atomic number).

e.g. O is [He]$2s^2$**$2p^4$**, O^- is [He]$2s^2$**$2p^5$** & O^{2-} is [He]$2s^2$**$2p^6$**, or [Ne].

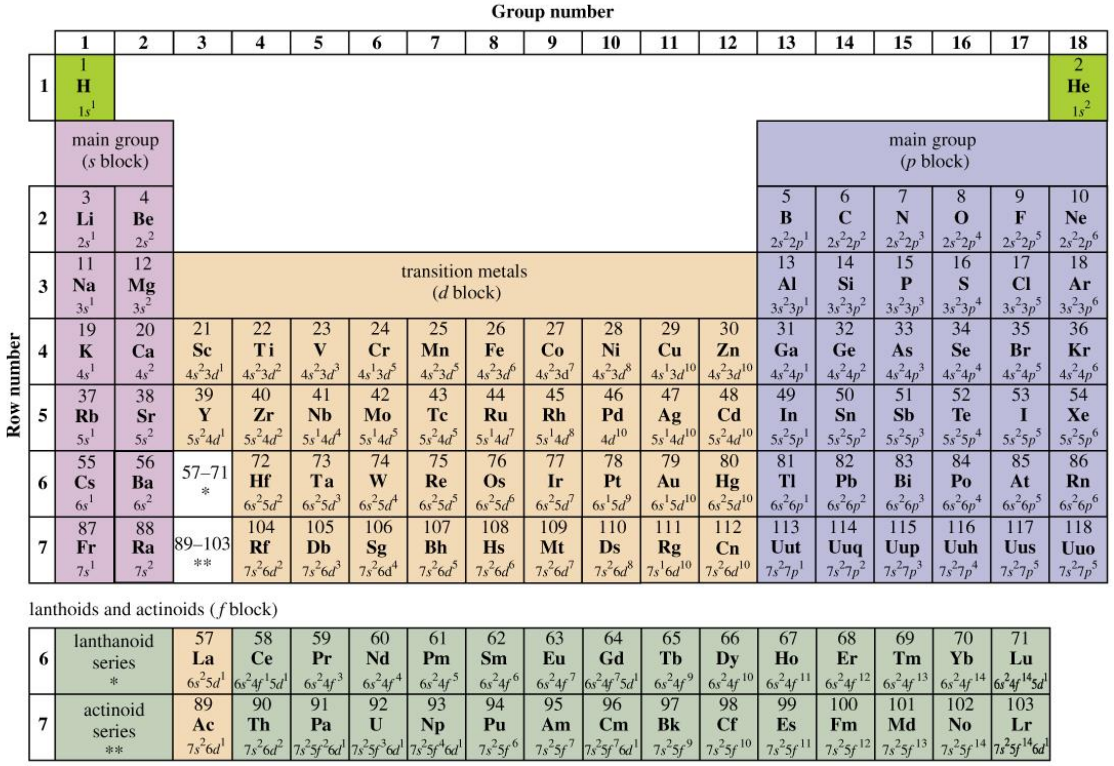
Cations

To determine the electronic configuration of a cation **remove the last electron that was added** (reverse order), **except for transition metal ions** where *s* orbital electrons are removed first.

ION CONFIGURATION

What is the ground state ion with the electronic configuration isoelectronic with argon?

- **A** F^-
- **B** Ne
- **C** Cl^-
- **D** Ca^+
- **E** Don't know



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ION CONFIGURATION

What is the electronic configuration of Mg^{2+}?

A [Ne]

B $[Ne]3s^2$

C $[Ne]3s^{2+}$

D $[Ne]3s^2 3p^2$

E Don't know

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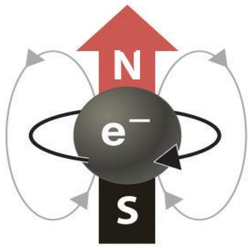
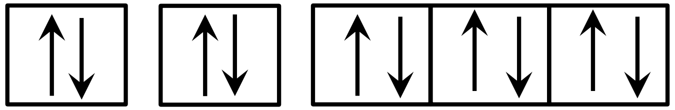
MAGNETISM



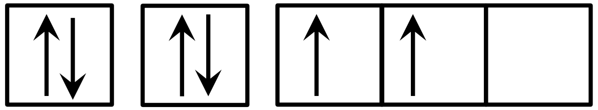
Electron aligned with magnetic field:

Paired electrons (in an orbital) are repelled very weakly by a magnetic field;

atoms where all electrons are paired are **diamagnetic.**



Unpaired electrons are attracted more strongly into the field. Atoms with unpaired electrons **paramagnetic**. Paramagnetism arises from the spin component of the electron that generates its own magnetic field.



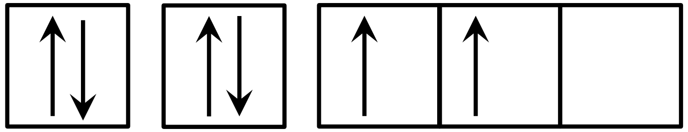
An atom containing many paired electrons, but only one unpaired electron will have a net attraction into the field, as the **paramagnetic** effect (attraction) of just one electron is stronger than the total **diamagnetism** (repulsion) of many electrons.

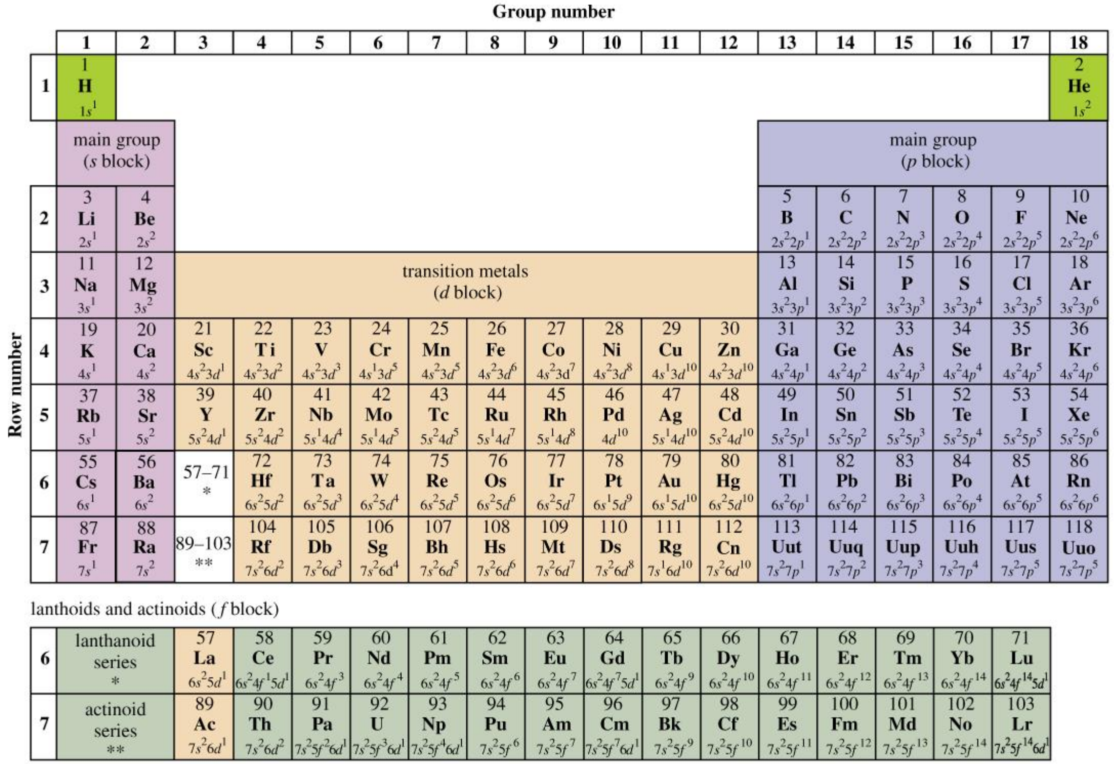
Any substance composed of atoms or molecules containing unpaired electrons will be paramagnetic.

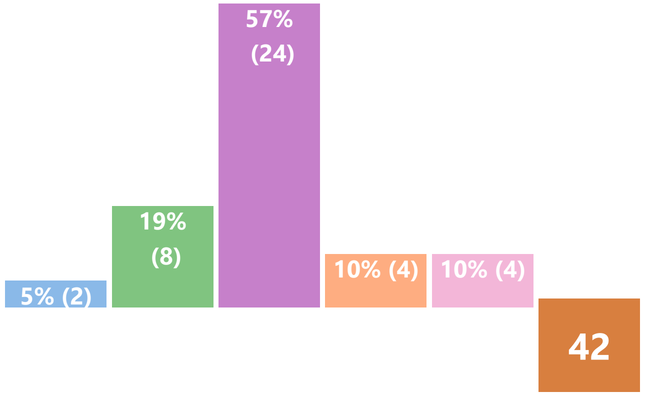
MAGNETISM

Which atom is paramagnetic?

- A He
- B Be
- C C
- D Ne
- E Don't know







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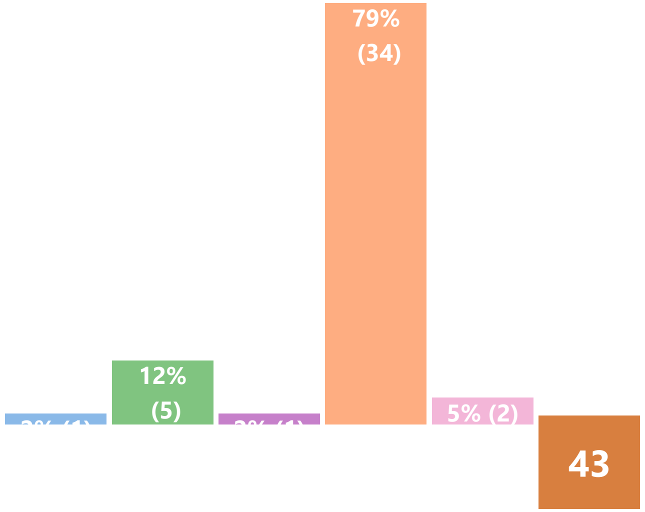
MAGNETISM

Which ion is paramagnetic?

- **A** Li^{+}
- **B** O^{2-}
- **C** F^{-}
- **D** Mg^{+} ✓
- **E** Don't know



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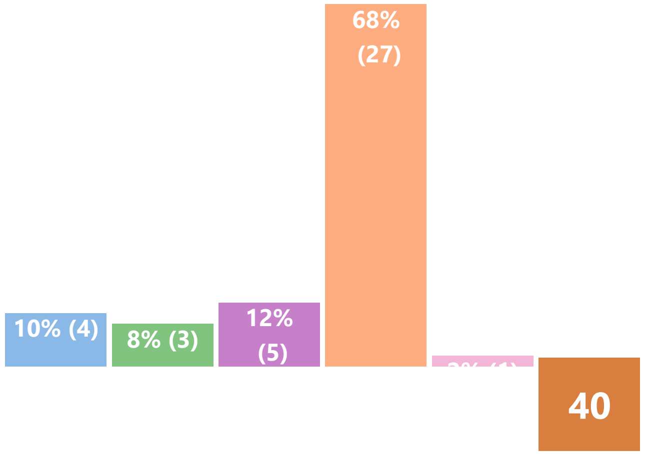
MAGNETISM

Which transition metal is **diamagnetic**?

A	Sc
B	Cr
C	Cu
D	Zn
E	Don't know



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LEARNING OUTCOMES

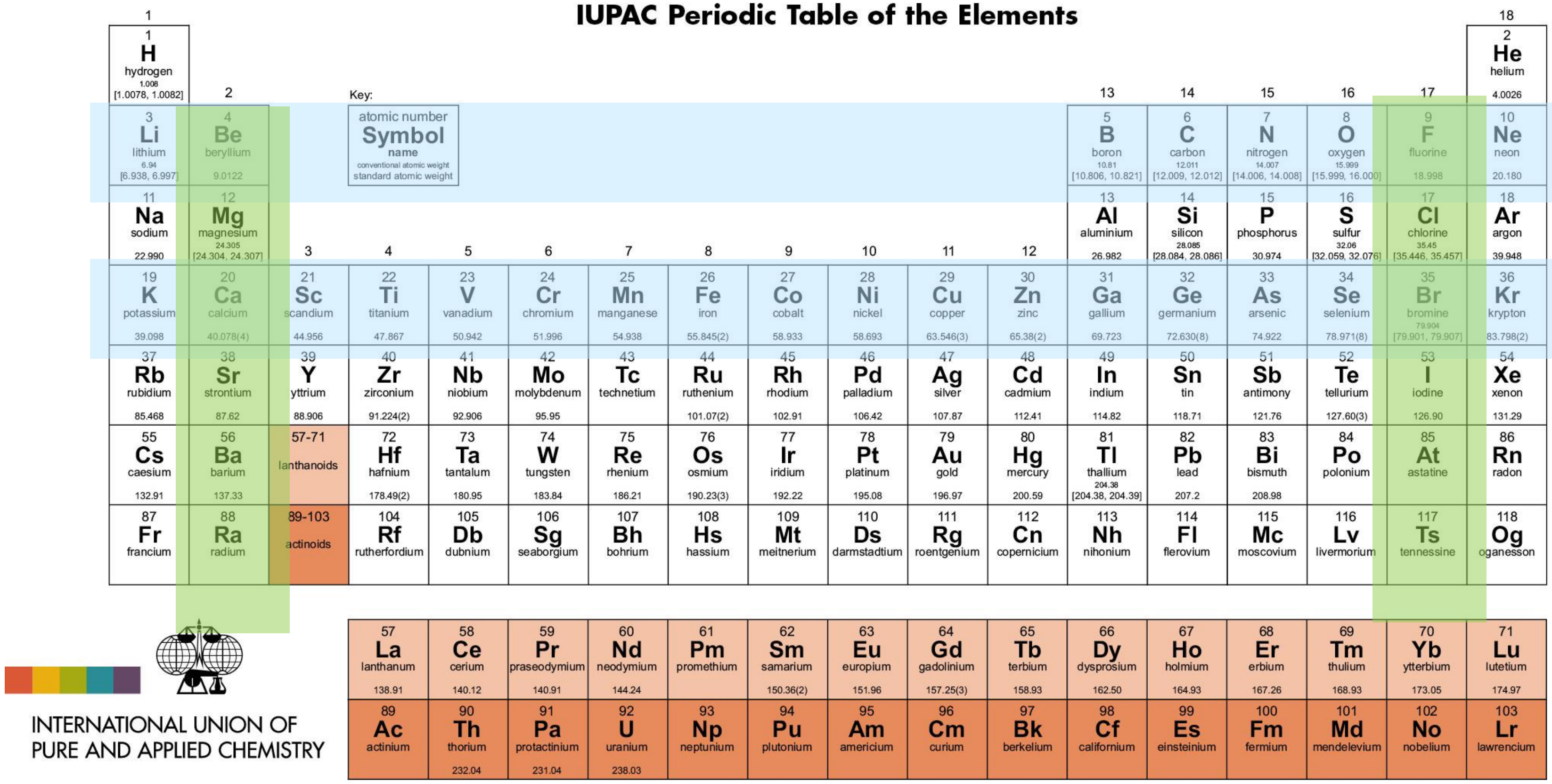
- Identify isoelectronic species and predict relative sizes of these species.
- Predict relative sizes, ionization energies, electron affinities and electronegativities of atoms based on position in the periodic table.

PERIODICITY

Predicting properties of atoms, ions and bonding based on where the elements are on the periodic table.

Period: (row) going *across* the periodic table

Group: (column) going down the periodic table

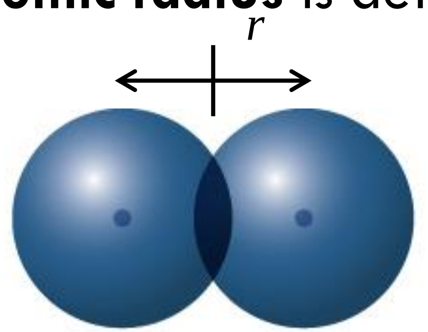


For notes and updates to this table, see www.iupac.org. This version is dated 28 November 2016.

ATOMIC SIZE

How do we define the size of an atom? Remember electrons exist in a "cloud" surrounding the nucleus and as such the 'edge' of the atom is difficult to define. There are a number of different methods to describe the radius of an atom such as van der Waals radius, ionic radius or covalent radius.

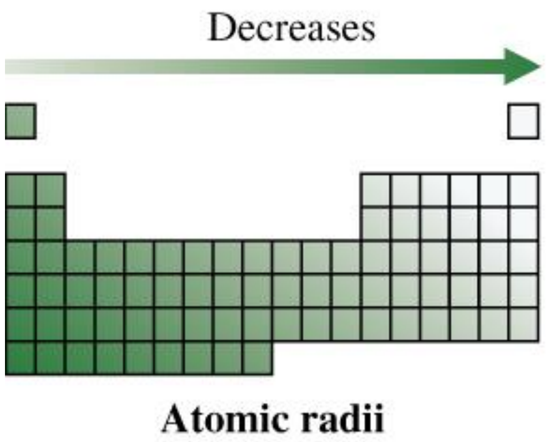
For elements that exist as single atoms in the solid state (metals, noble gases) the **atomic radius** is defined as one **half the distance between adjacent nuclei.**



For diatomic molecules (e.g. O_2, Cl_2) the **covalent radius** is defined as **half the internuclear distance.**

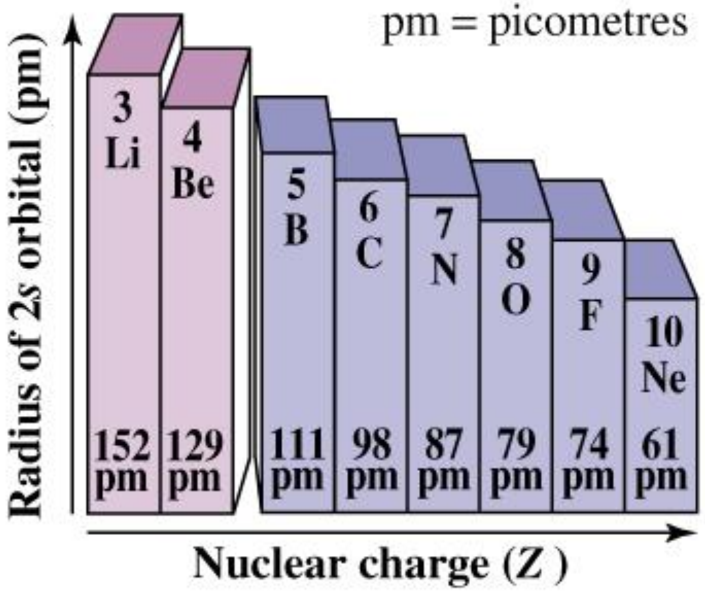
Note: the electron clouds could overlap to some extent

ATOMIC SIZE



Across a period (from left to right →):

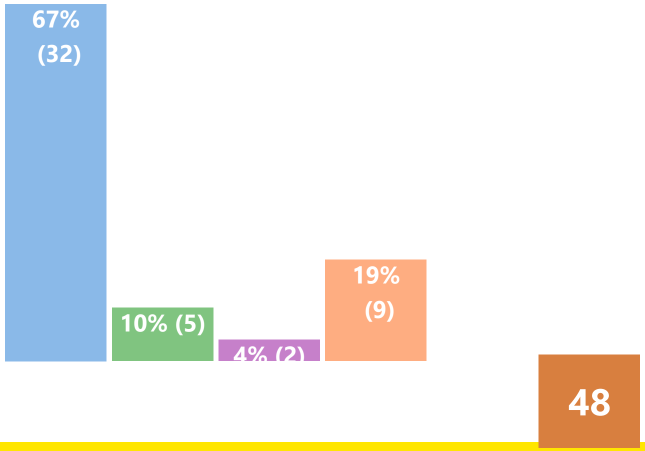
The atomic size *decreases*.



ATOMIC SIZE

Why do you think atomic size *decrease across a period*?

- **A** There is an increase in nuclear charge
- **B** There are more electrons so they repel each other further away from the nucleus
- **C** The shell is increasing in size
- **D** Shielding from inner electrons means outer electrons are not as strongly attracted to the nucleus
- **E** Don't know



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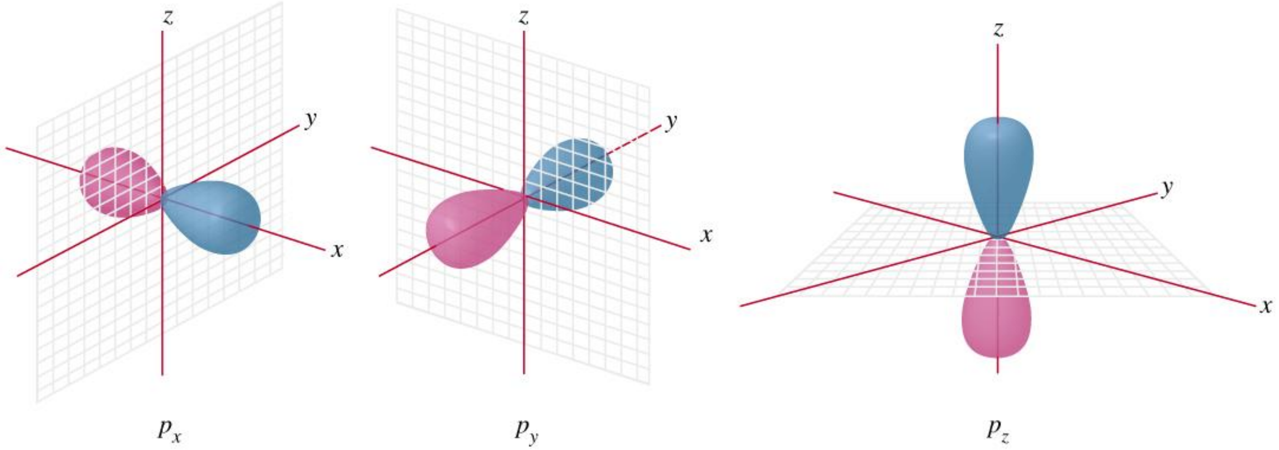
SHIELDING

Shielding is a measure of the positive charge experienced by an electron, taking into account the shielding of all the other electrons.

Remember that not all shielding is equal. **“Inner” electrons (core, lower *n*) shield “outer” electrons (valence, higher *n*).** The degree of shielding depends on the type of orbitals.

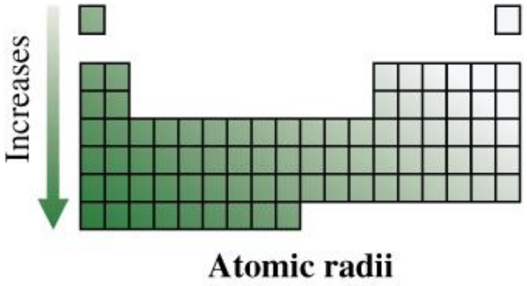


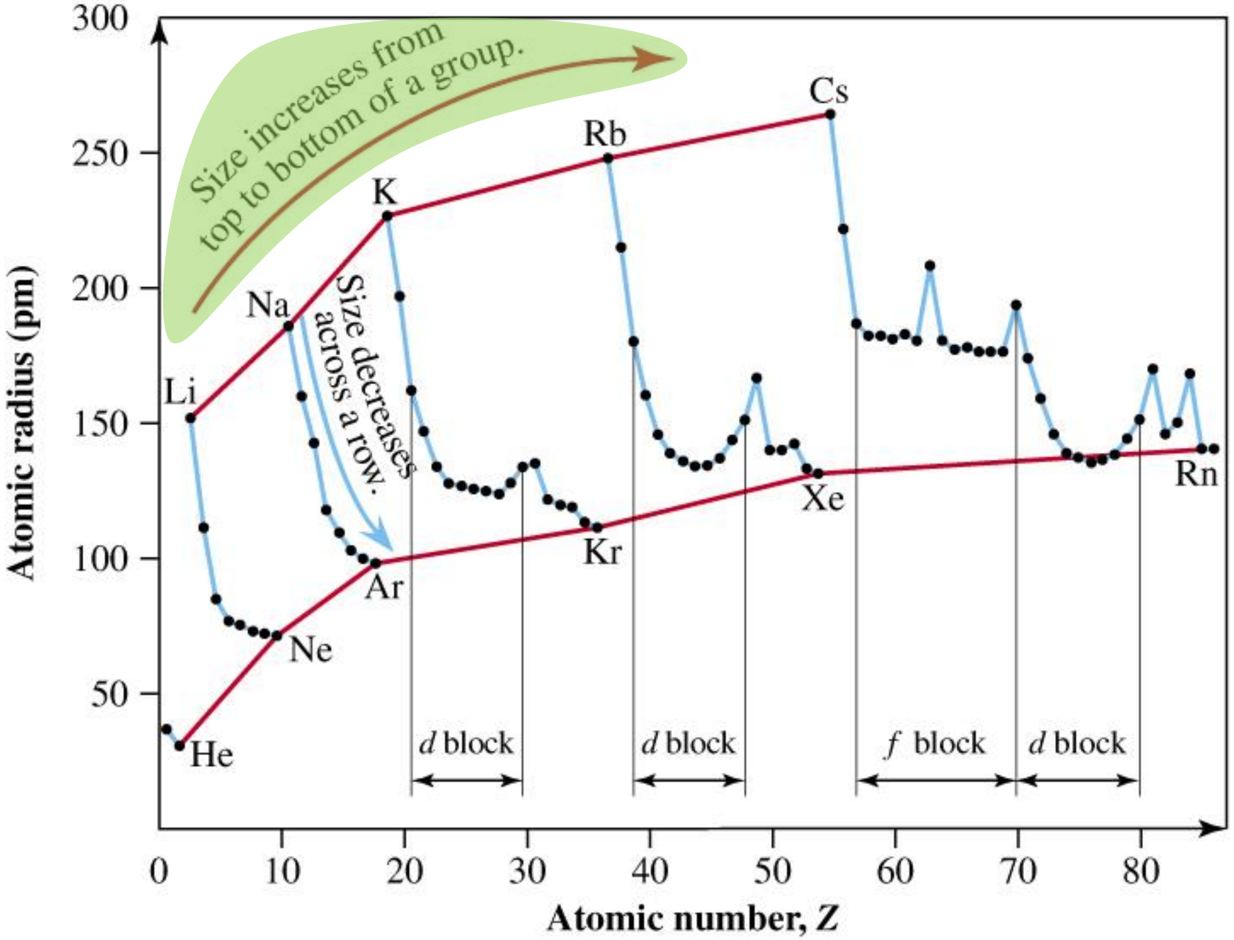
Within a shell (i.e. same *n*) the orbitals shield each other very poorly.

ATOMIC SIZE

Going down a group:

The atomic size *increases*





ATOMIC SIZE

Why do you think atomic size *increase down a group*?

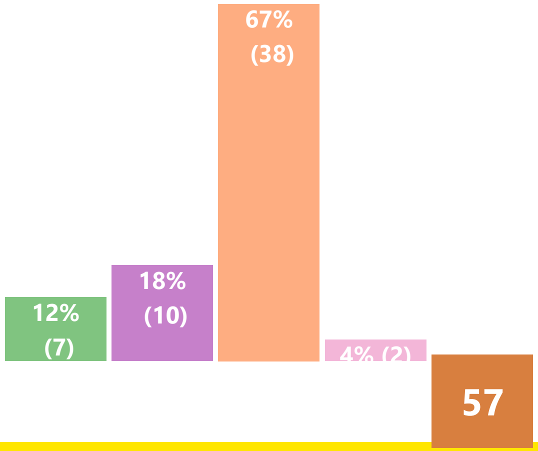
A There is an increase in nuclear charge — True, but opposite effect.

B There are more electrons so they repel each other further away from the nucleus — True

C The shell is increasing in size — True

D Shielding from inner electrons means outer electrons are not as strongly attracted to the nucleus — The main reason

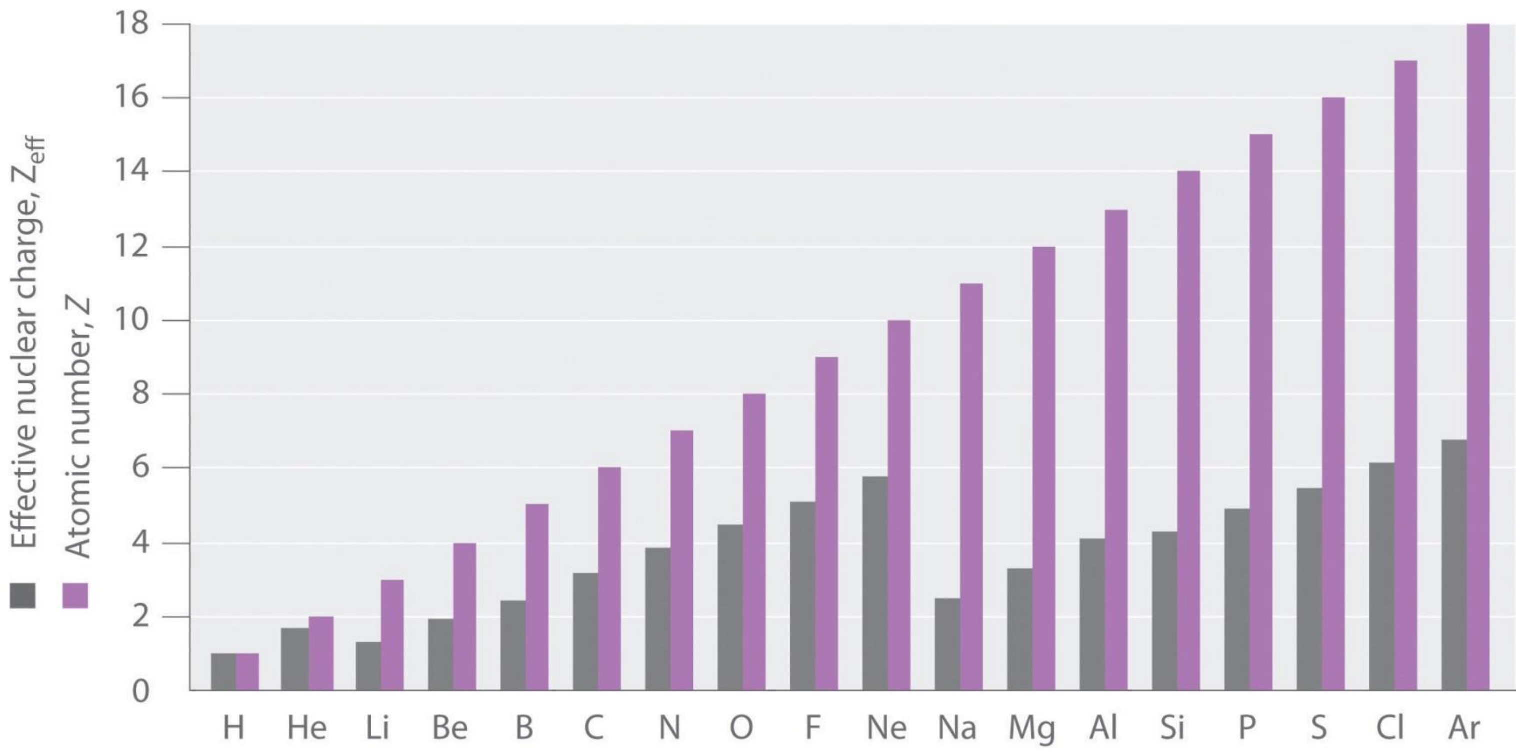
E Don't know



EFFECTIVE NUCLEAR CHARGE

In chemistry everything is a competition – like for electrons!

Shielding from core electrons causes the **effective nuclear charge** (Z_{eff}) to be **much smaller** than the actual nuclear charge (Z).



Relationship between the Effective Nuclear Charge Z_{eff} and the Atomic Number Z for the Outer Electrons of the Elements of the First Three Rows of the Periodic Table

http://images.flatworldknowledge.com/averillfwk/averillfwk-fig06_028.jpg

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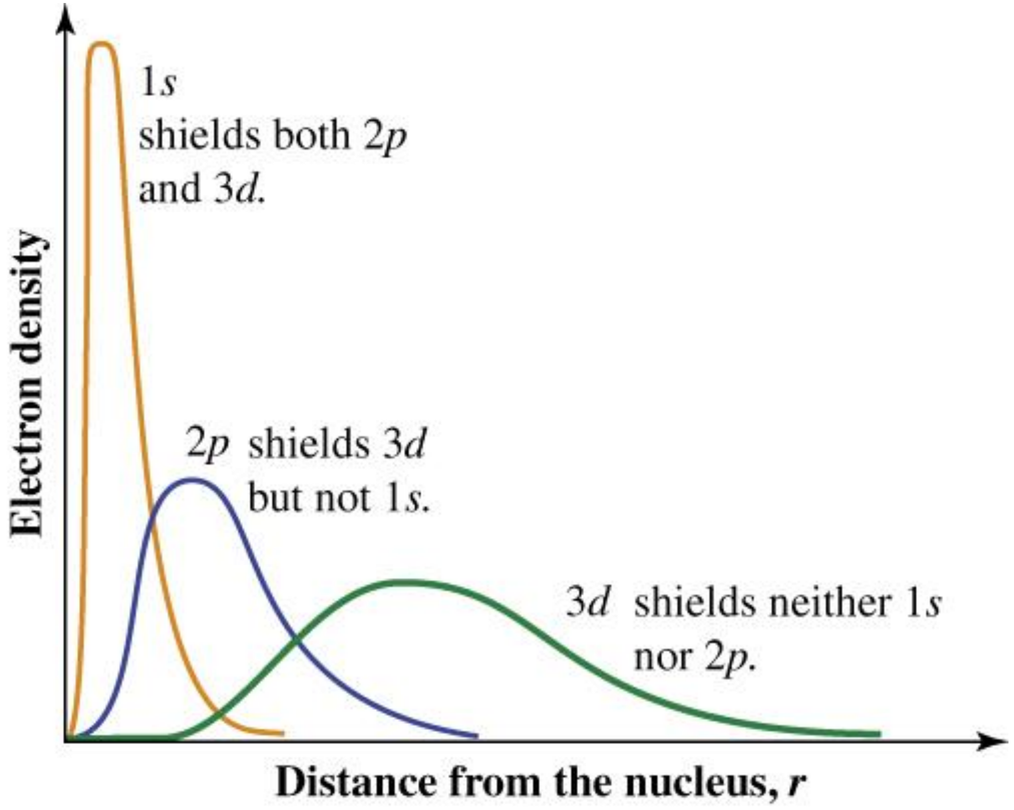
SHIELDING

Shielding is a measure of the positive charge experienced by an electron, taking into account the shielding of all the other electrons.

Remember that not all shielding is equal. **"Inner" electrons (core, lower *n*) shield "outer" electrons (valence, higher *n*).** The degree of shielding depends on the type of orbitals.



Within a shell (i.e. same *n*) the orbitals shield each other very poorly.

IONIC SIZE

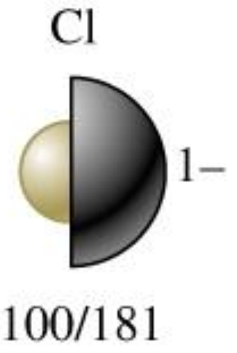
Cations

Smaller than their parent atom

Atomic/ionic radius

Anions

Larger than their parent atom



Atomic/ionic radius

IONIC SIZE

Why are cations smaller than their parent atom?

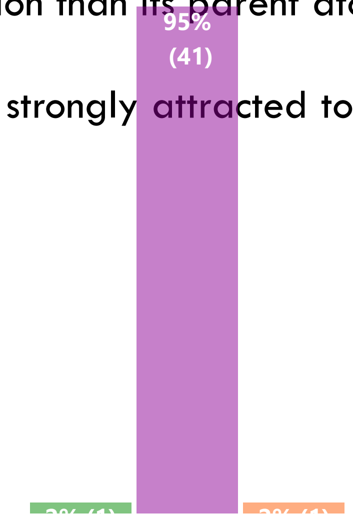
A There is an increase in nuclear charge

B There are more electrons so they repel each other less

C The effective nuclear charge is greater (per electron) for a cation than its parent atom

D Shielding from inner electrons means outer electrons are not as strongly attracted to the nucleus

E Don’t know



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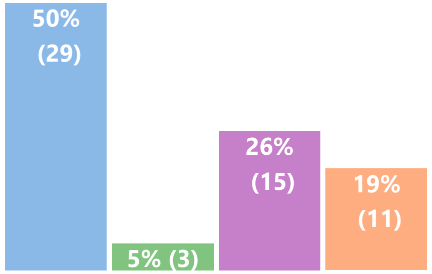
IONIC SIZE

Which ion is largest? *Hint: they are all isoelectronic.*

- **A** O^{2-}
- **B** F^{-}
- **C** Na^{+}
- **D** Mg^{2+}
- **E** Don't know



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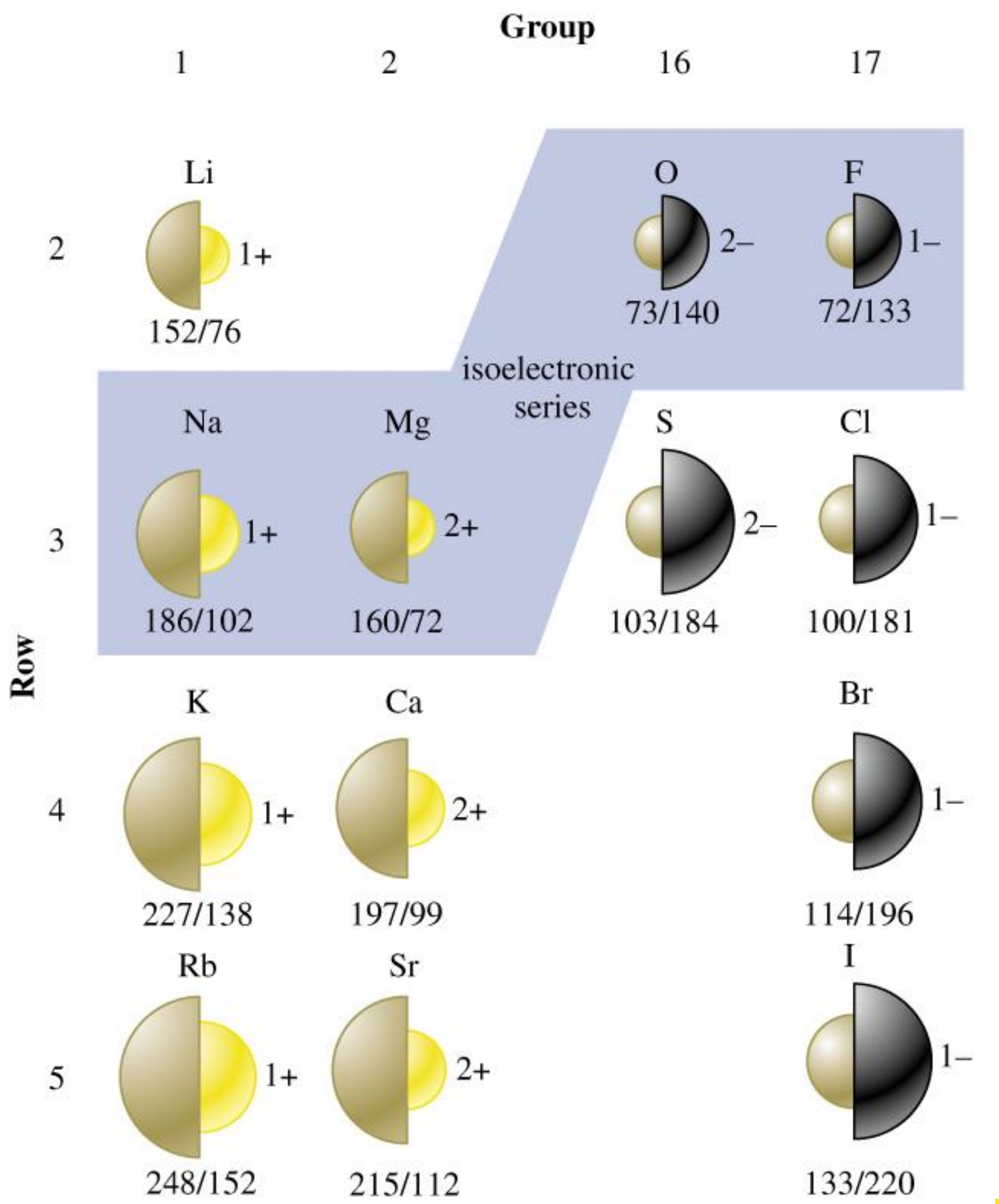
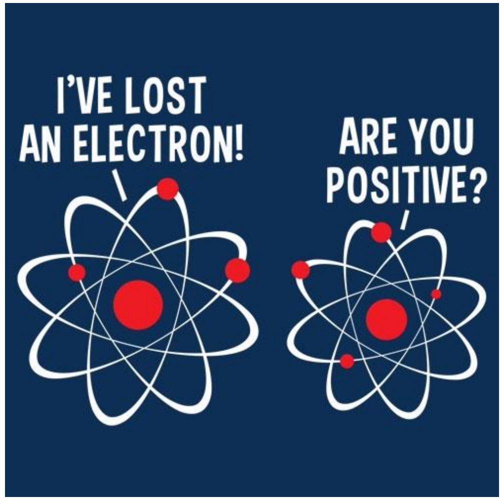
IONIC SIZE

The greater the charge the greater the effect.



FIRST IONISATION ENERGY



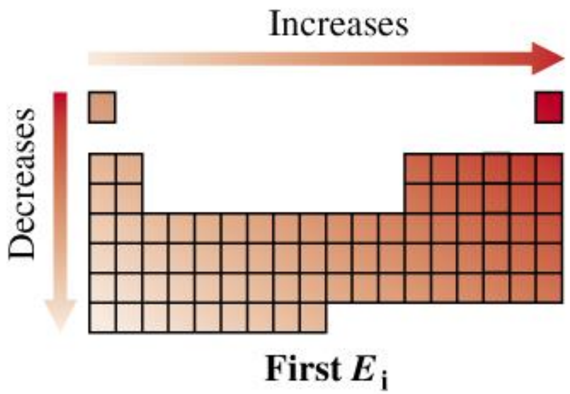
Ionisation energy: the energy required to remove an electron from an atom (measured in the gas phase).

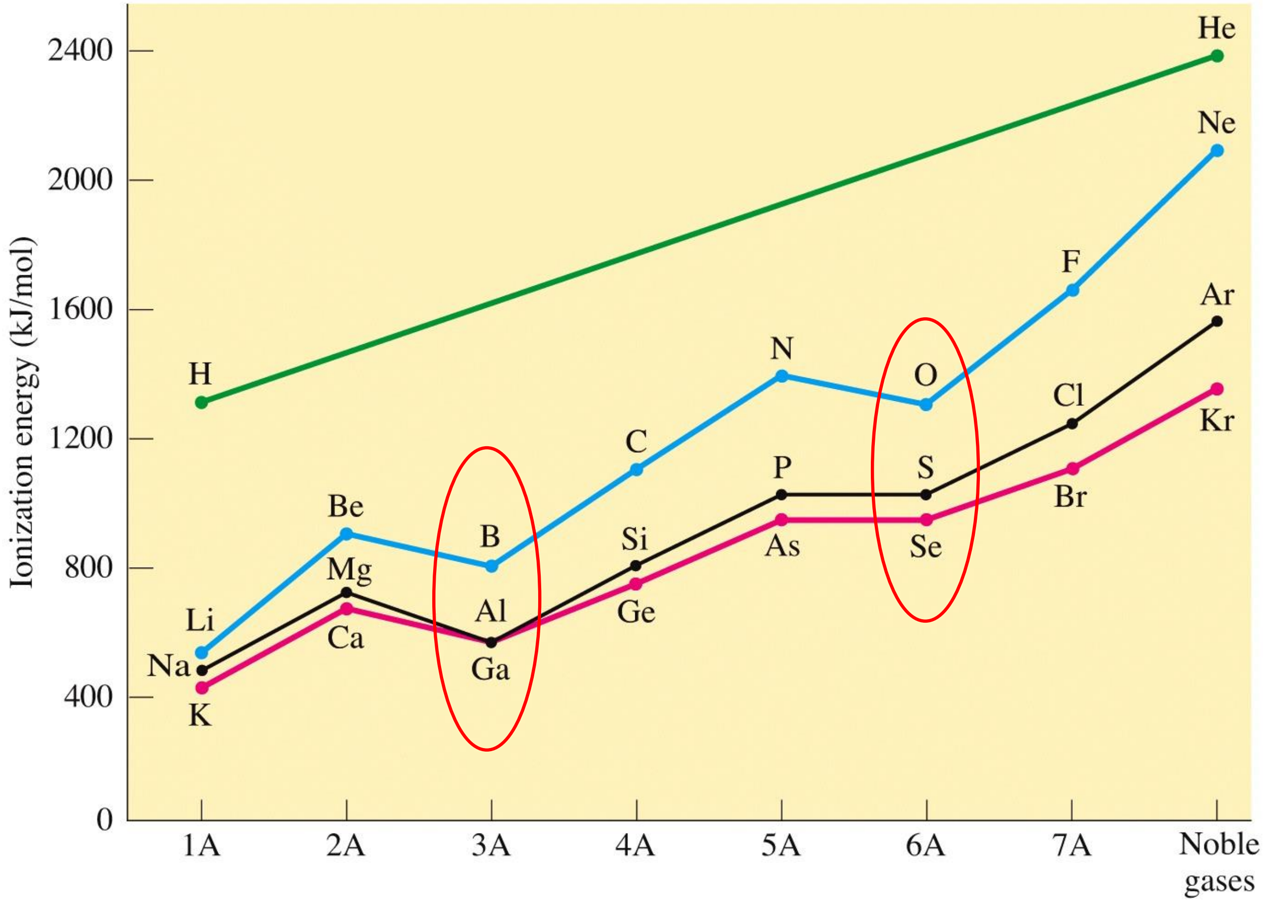
$$\text{atom (g)} \rightarrow \text{ion}^{+}\text{(g)} + \text{e}^{-}\text{(g)}$$

The **first ionisation energy corresponds to the first electron being removed from a neutral atom.** The second ionisation energy is the energy required to remove the next electron (i.e. forming a cation with 2+ charge).

Successive ionisations require more energy to remove subsequent electrons, as the remaining electrons are attracted more strongly to the nuclear charge of the resulting cation.

FIRST IONISATION ENERGY

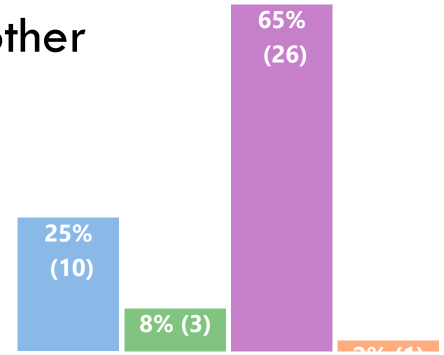




IONISATION ENERGY

Why does ionisation energy increase across a period?

- **A** Atoms get closer to noble gas configuration
- **B** Electrons are further away from the nucleus
- **C** The nuclear charge increases
- **D** There are more electrons so they repel each other
- **E** Don't know



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IONISATION ENERGY

Why does ionisation energy decrease down a group?

- **A** Atoms get closer to noble gas configuration
- ✓ **B** Electrons are further away from the nucleus
- **C** The nuclear charge increases
- **D** There are more electrons so they repel each other
- **E** Don’t know

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FIRST IONISATION ENERGY