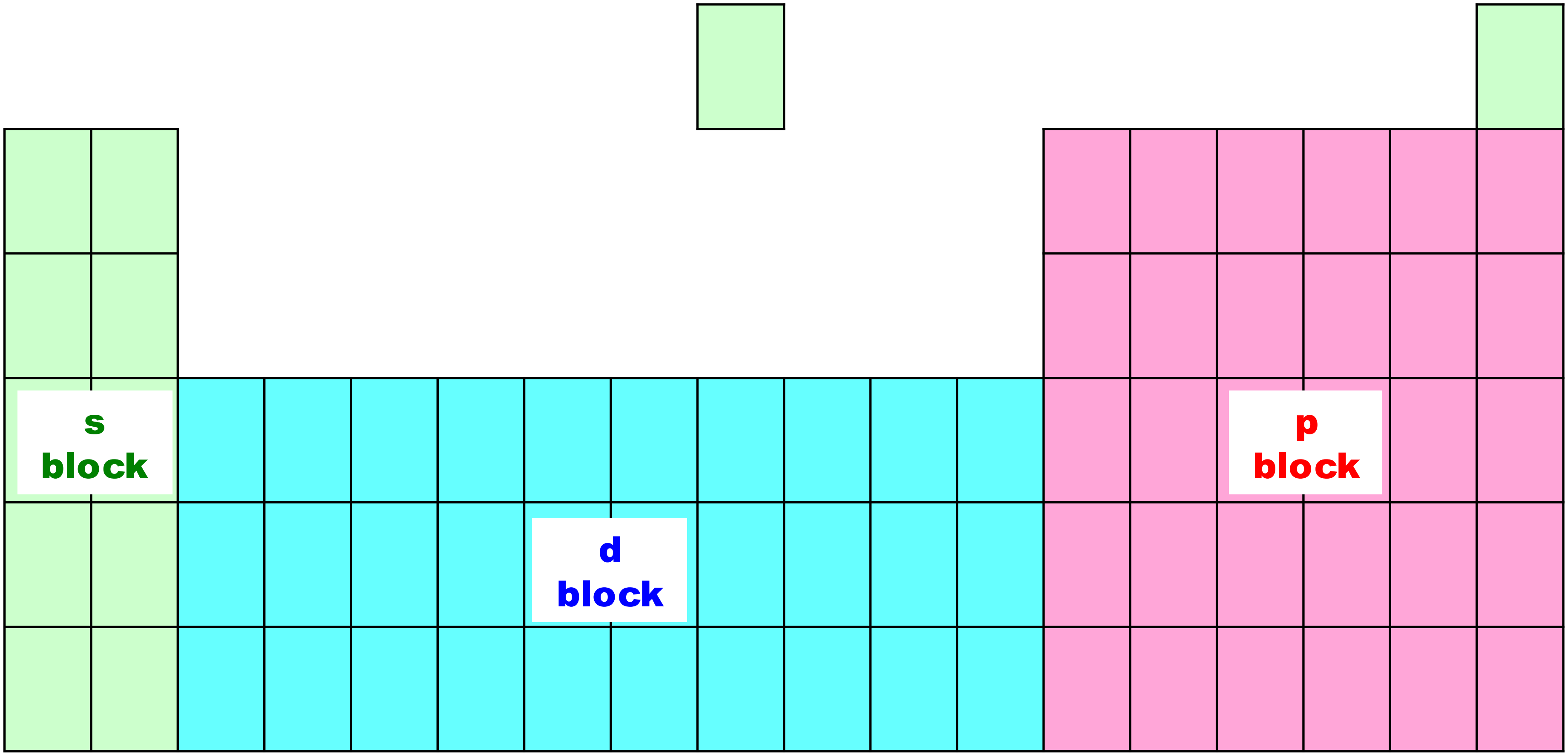
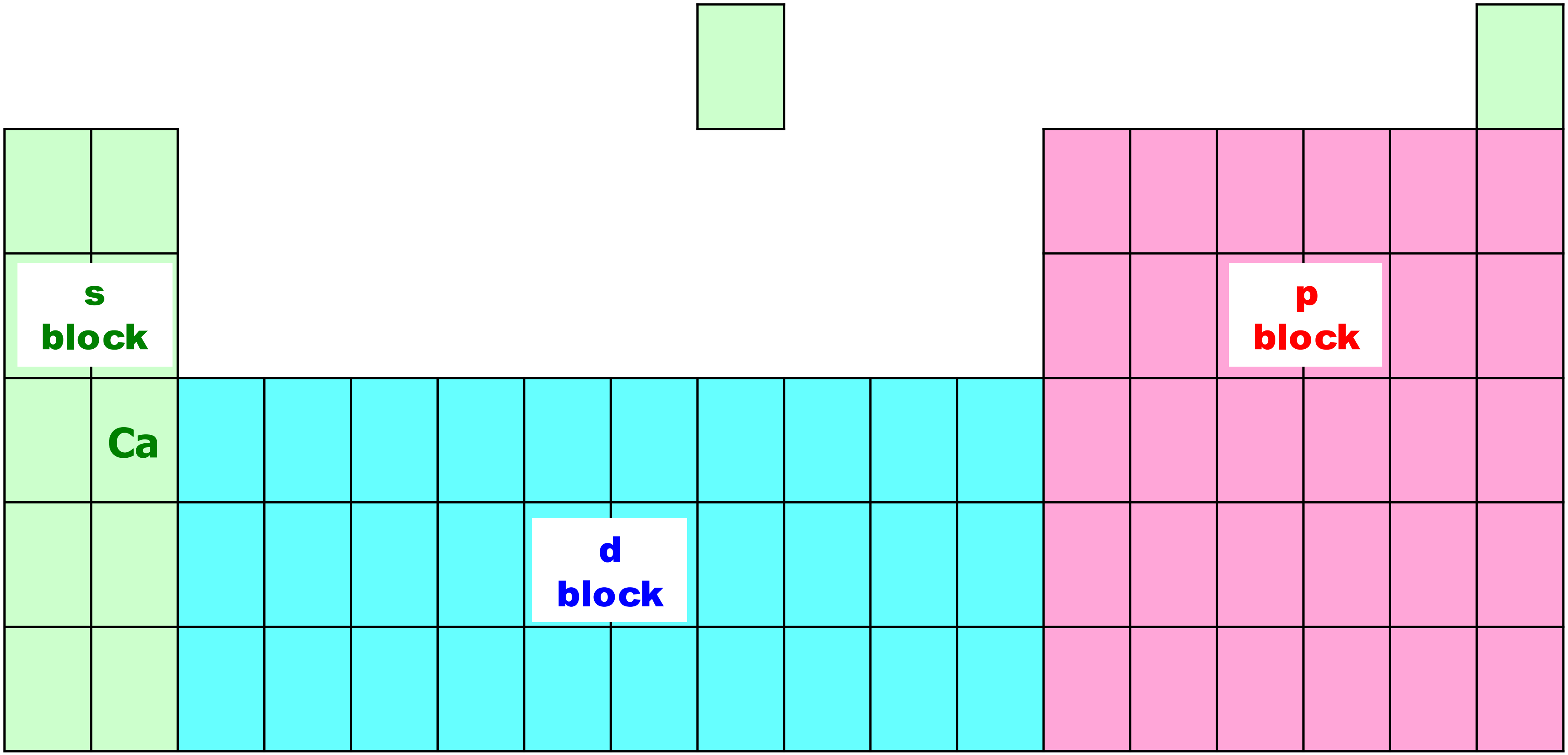
WWW.CHEMSHEETS.CO.UK

PERIODICITY



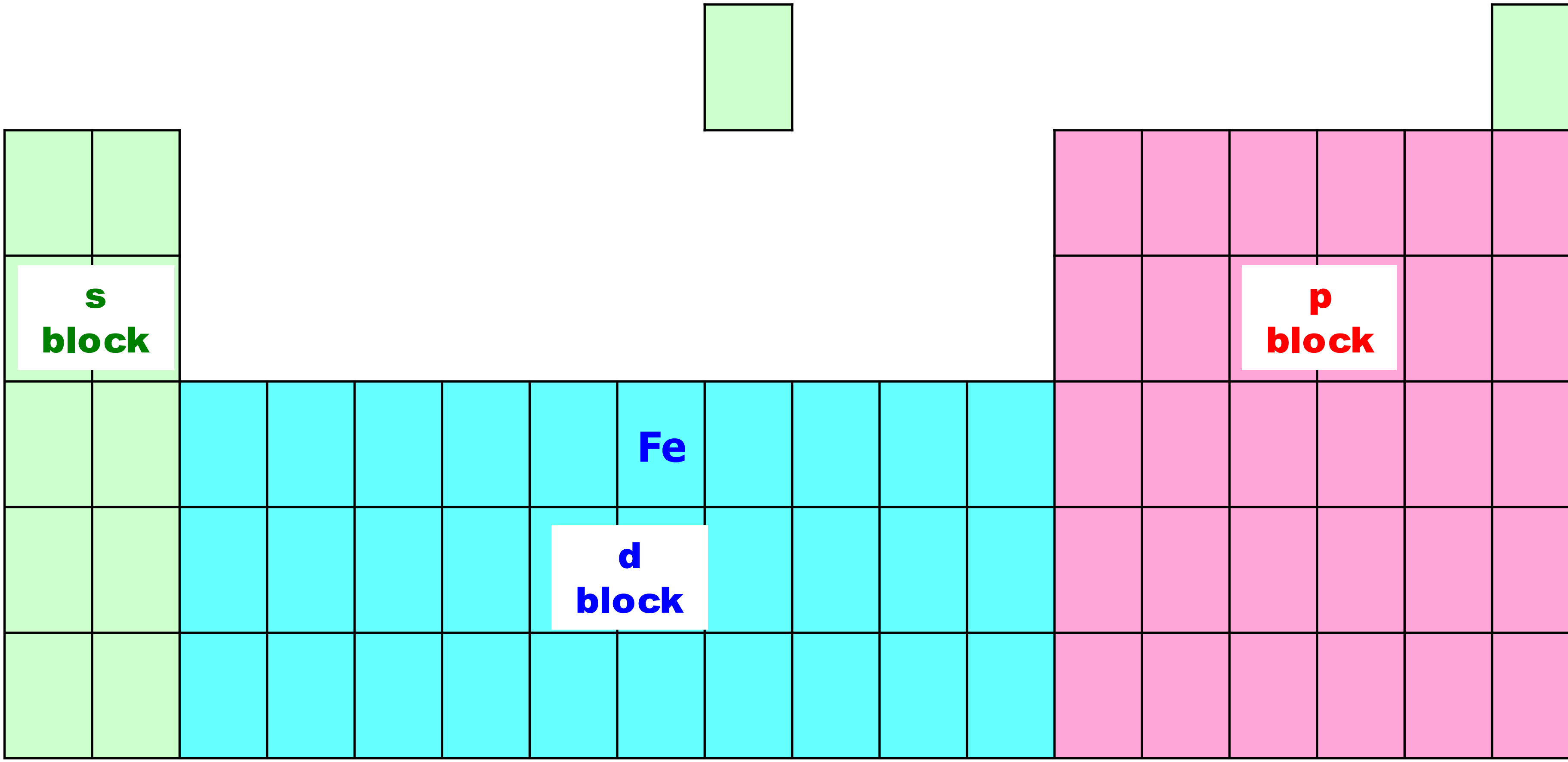
The orbitals that the highest energy electrons are in

Calcium 20 e^- $1s^2$ $2s^2$ $2p^6$ $3s^2$ $3p^6$ $4s^2$ **s block**



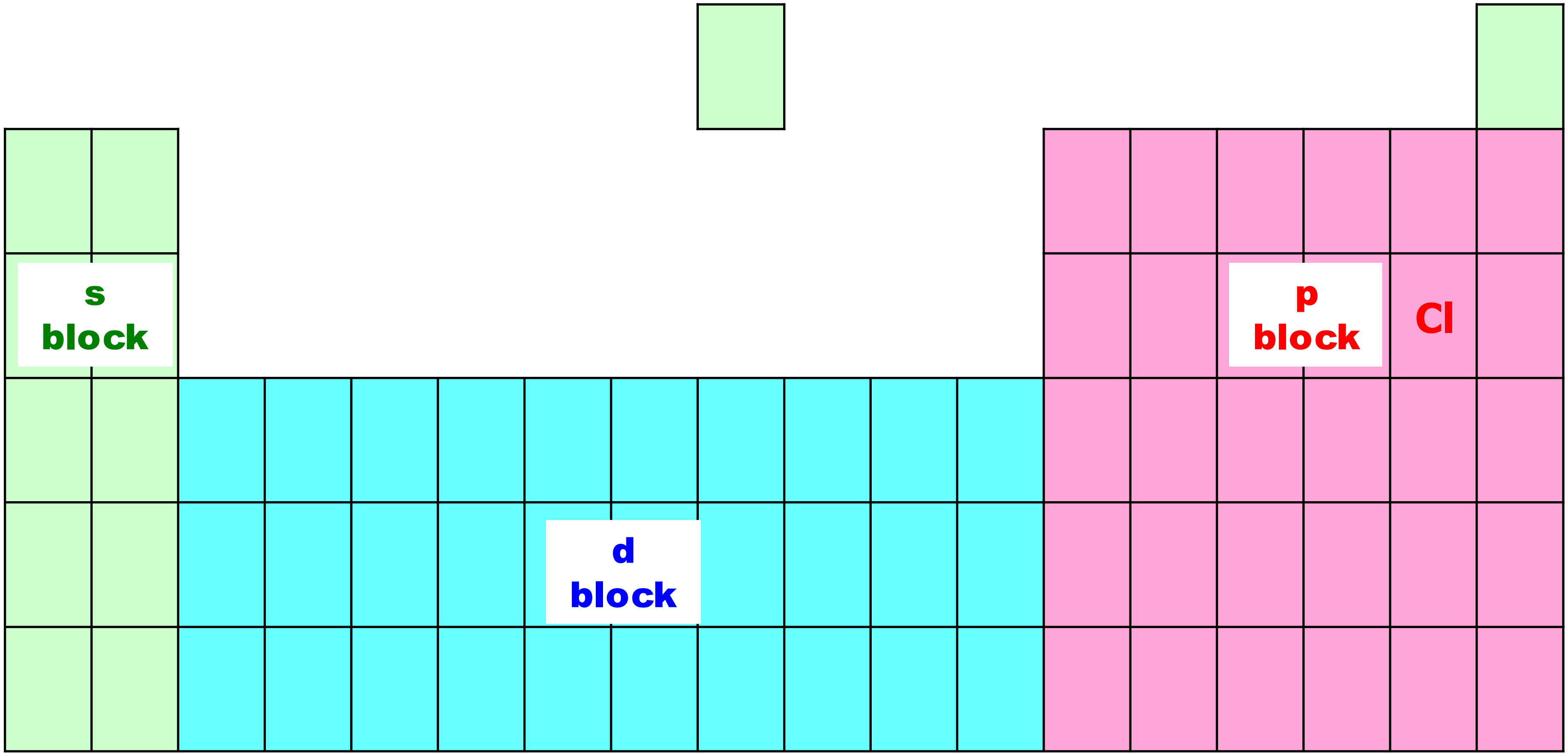
The orbitals that the highest energy electrons are in

Iron 26 e^- $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^6$ **d block**



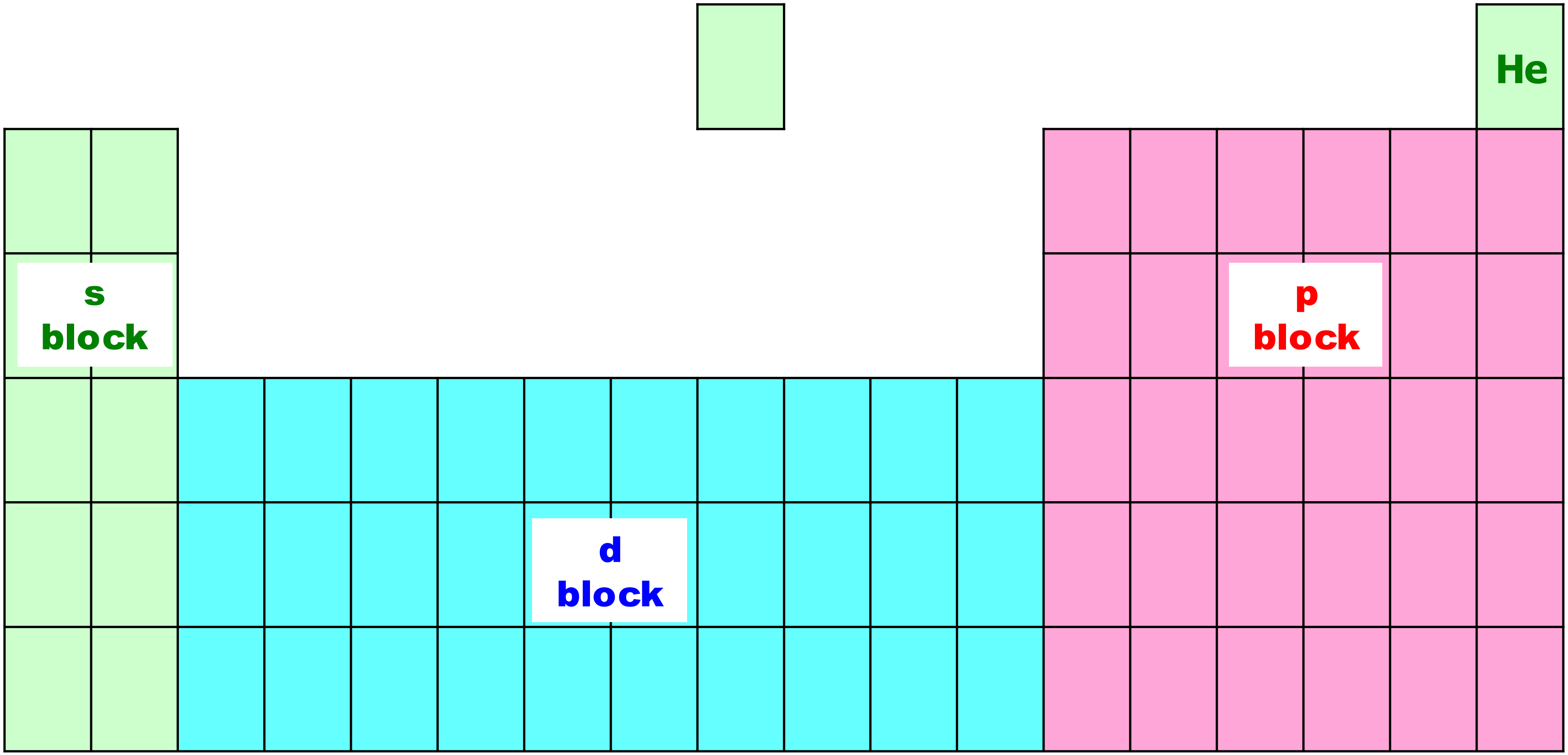
The orbitals that the highest energy electrons are in

Chlorine 17 e^- $1s^2\ 2s^2\ 2p^5$ **p block**



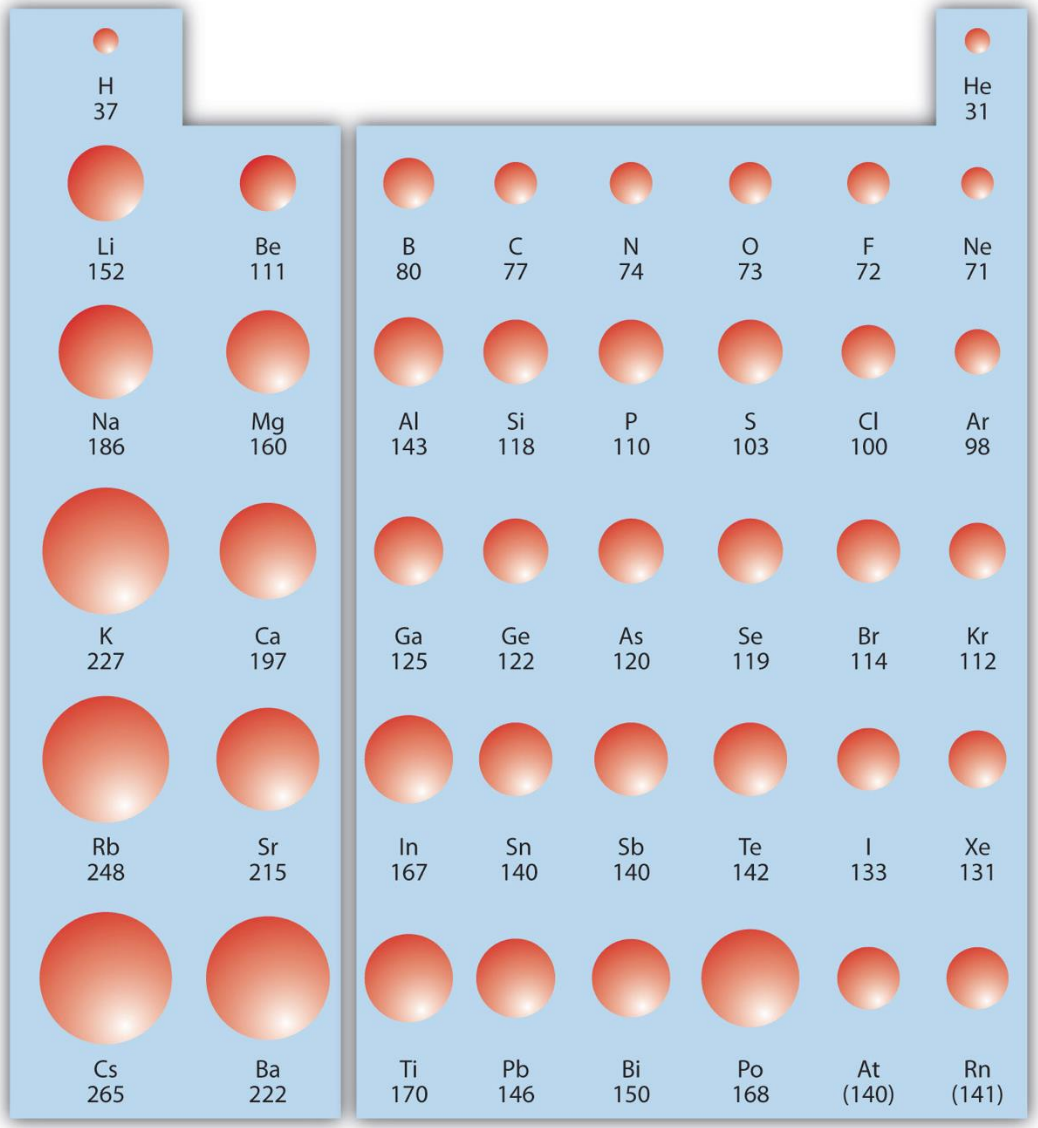
The orbitals that the highest energy electrons are in

Helium 2 e^- $1s^2$ **s block**



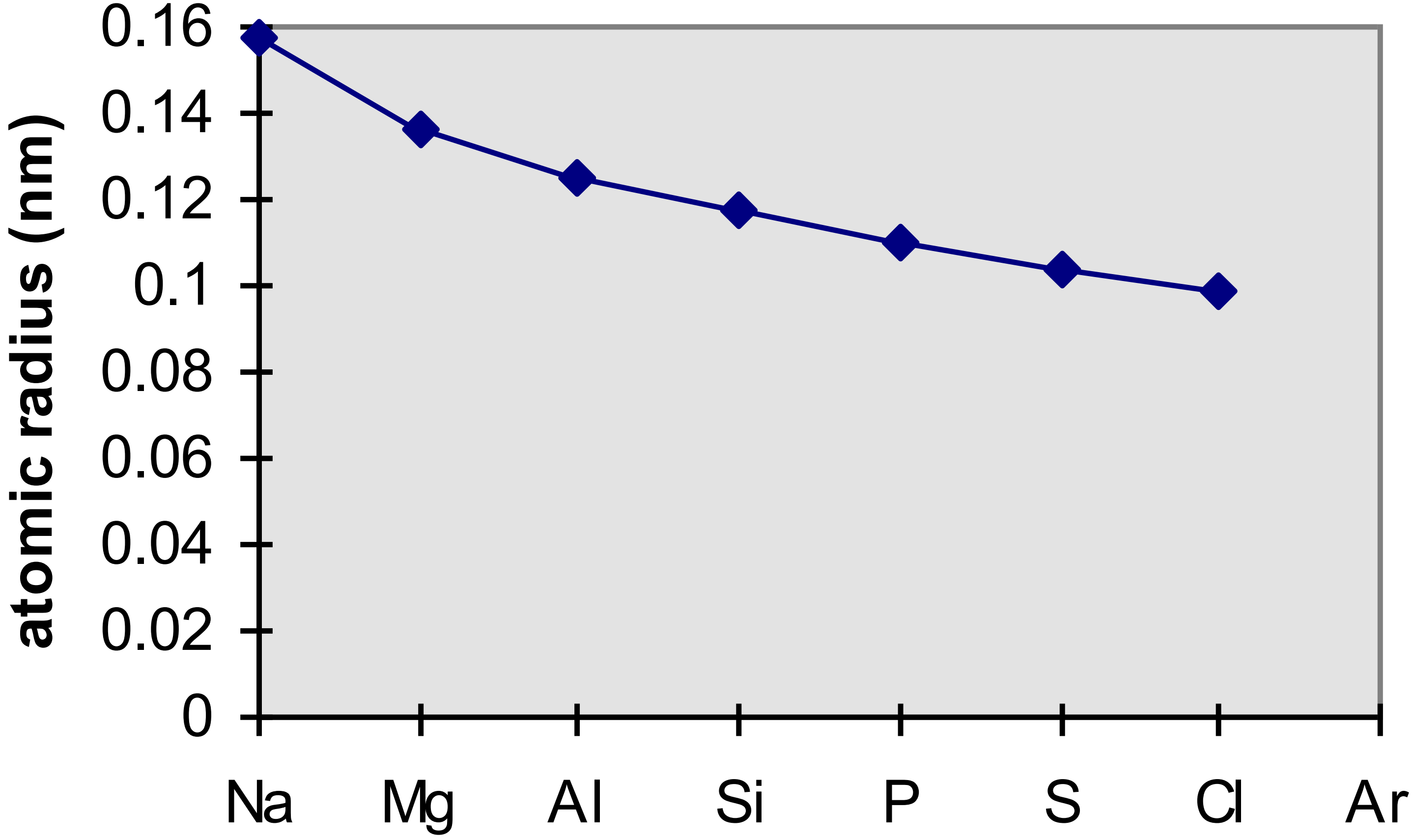
The orbitals that the highest energy electrons are in

ATOMIC RADIUS



From: http://2012books.lardbucket.org/books/beginning-chemistry/s12-05-periodic-trends.html

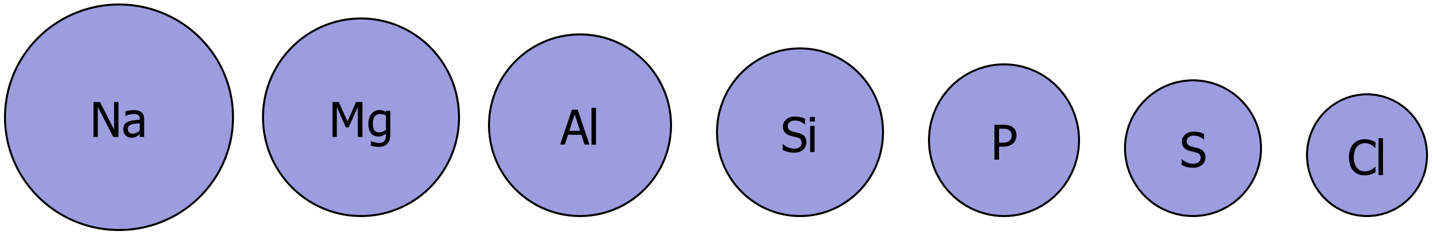
ATOMIC RADIUS



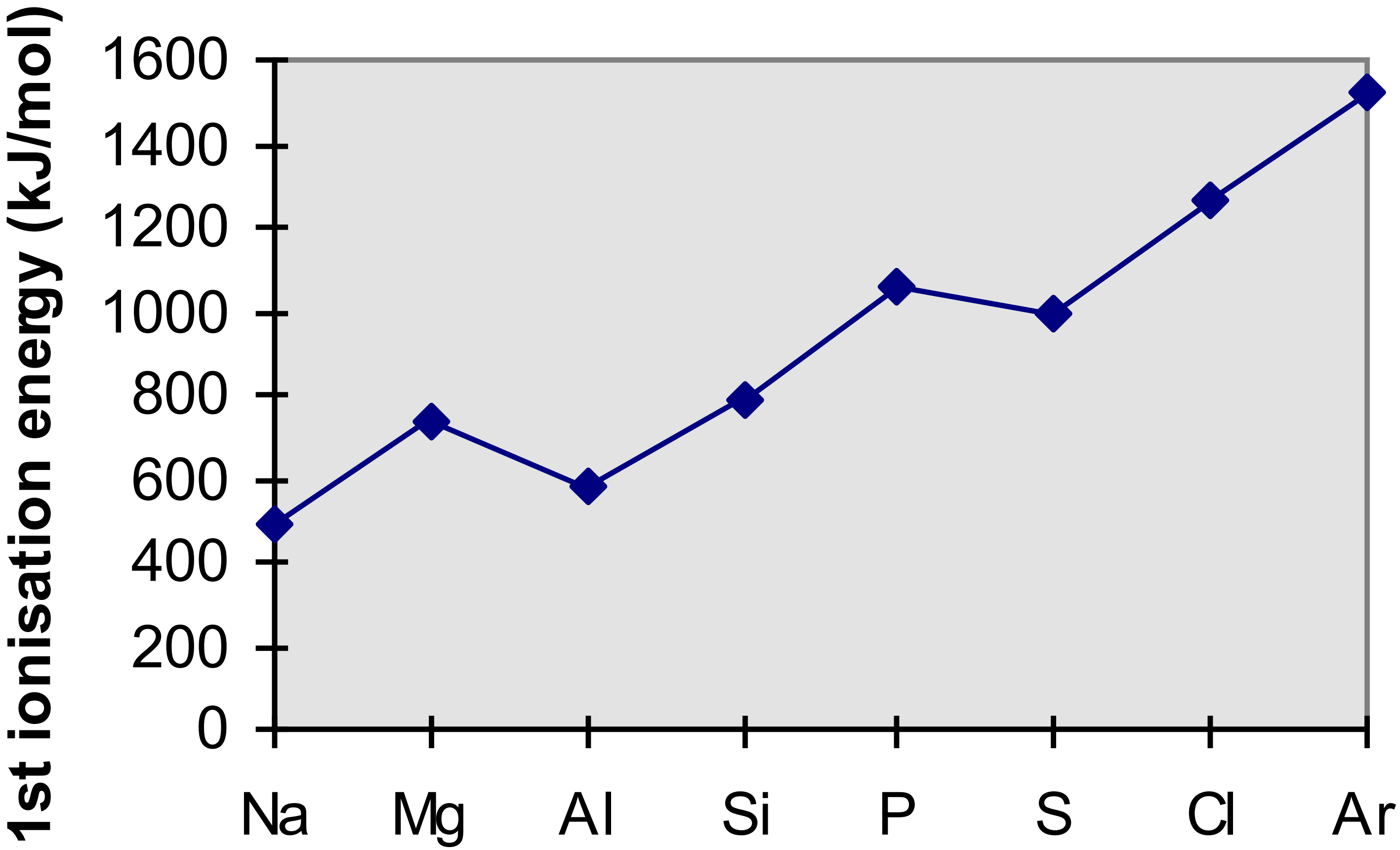
ATOMIC RADIUS

Across a period:

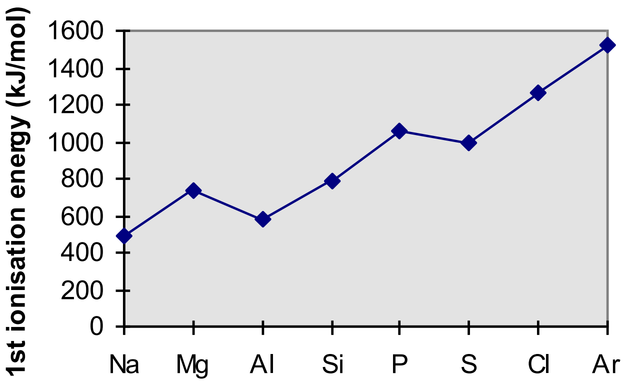
- outer electrons are in <u>same shell</u>
- more protons in nucleus
- same amount of shielding
- so stronger attraction between nucleus and outer shell electrons
- so outer shell electrons pulled closer to nucleus



1st IONISATION ENERGY



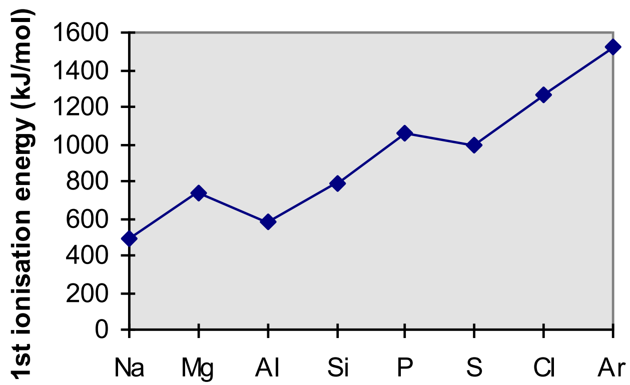
1st IONISATION ENERGY

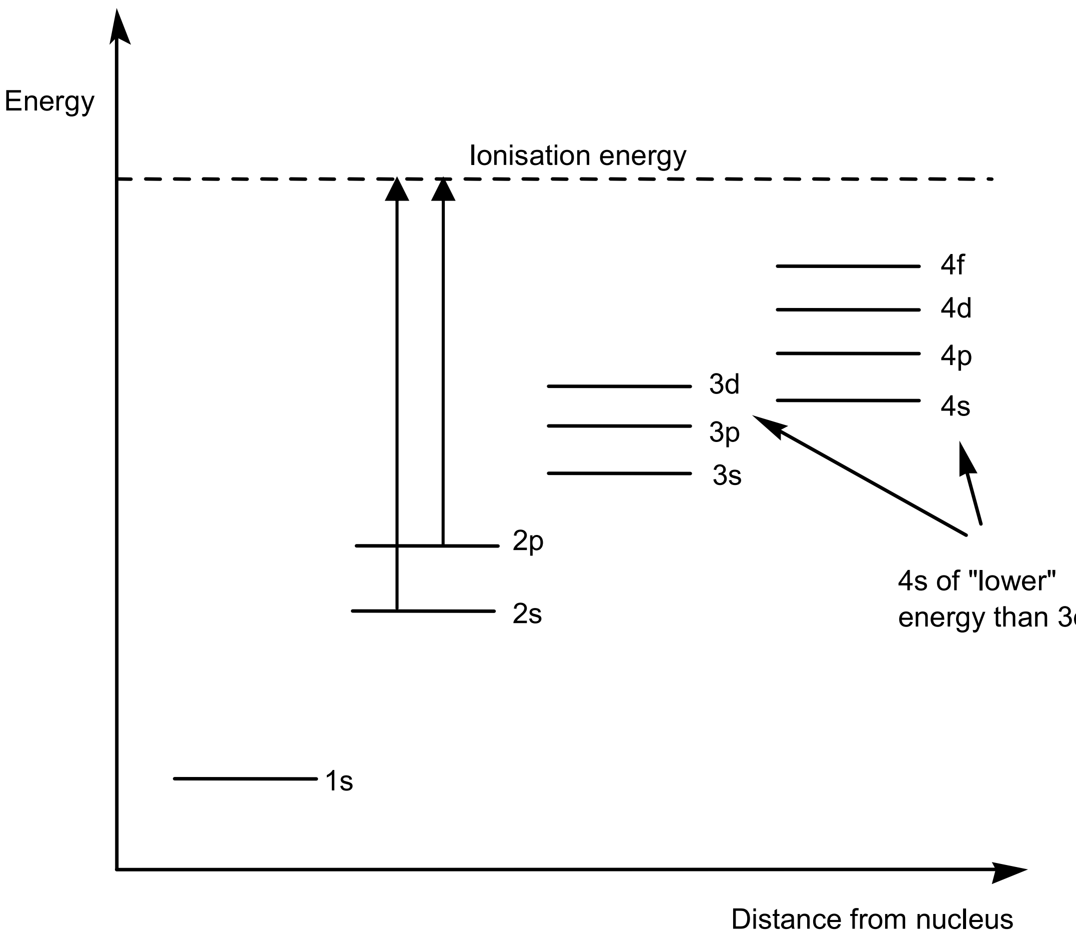


General trend across period

- more protons
- atoms get smaller
- therefore stronger attraction from nucleus to electron in outer shell

1st IONISATION ENERGY

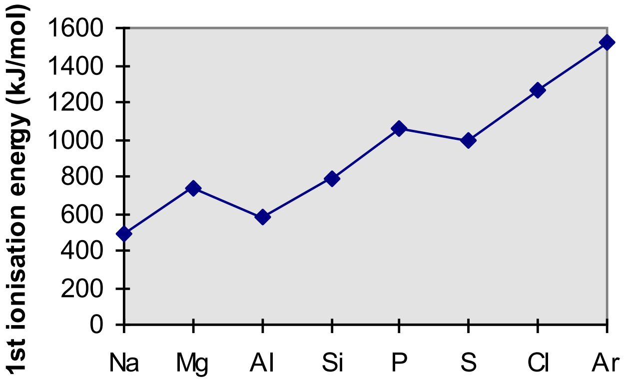




Group 2 → 3

- electron lost from:
  Group 3 = p orbital
  Group 2 = s orbital
- p orbital is higher energy than s orbital, so easier to lose electron.

1st IONISATION ENERGY



Group 5 → 6

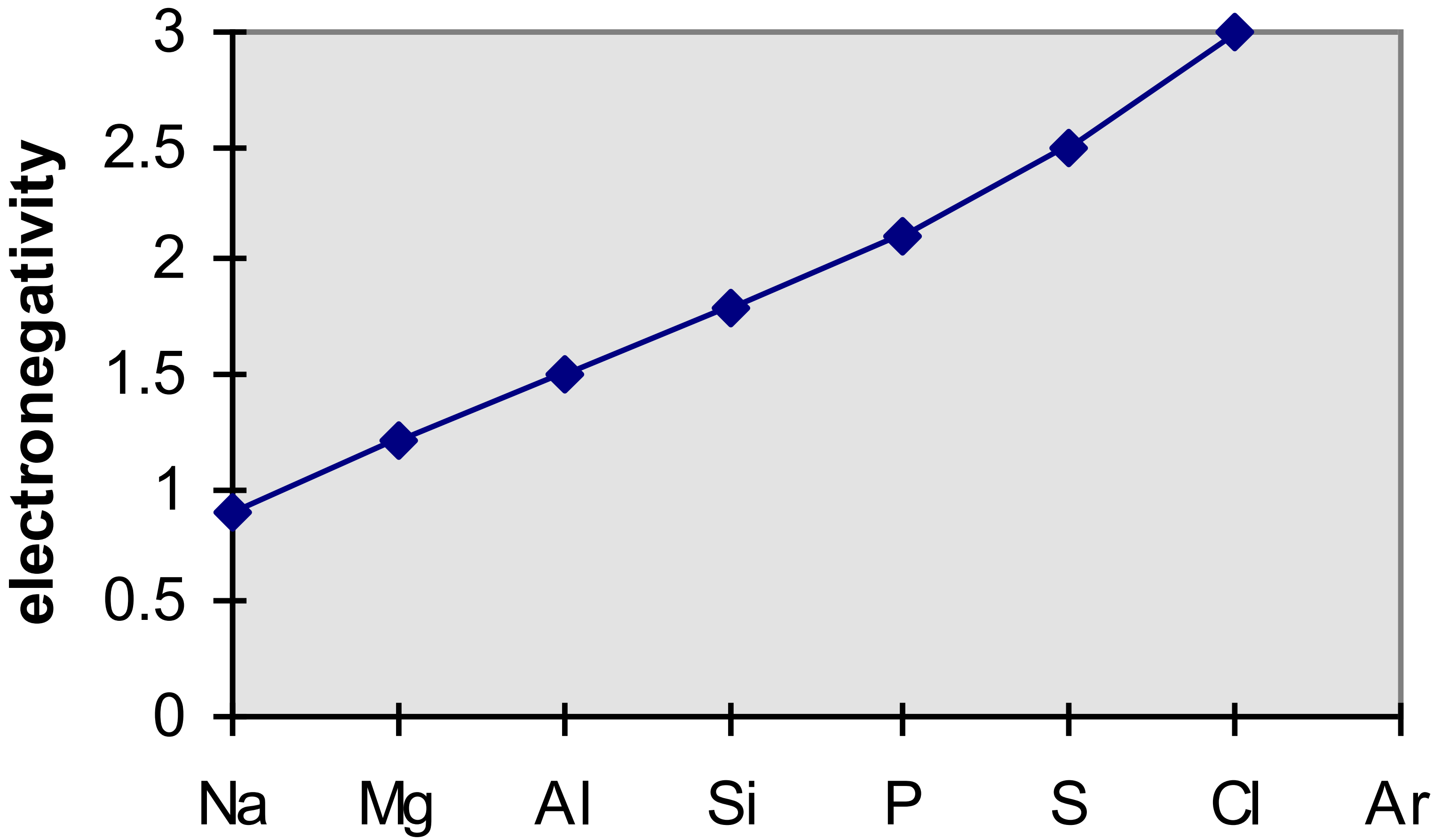
- Group 6 element loses electron from orbital with 2 electrons (p^4 ⍰⍰⍰).
- Group 5 element loses electron from orbital with 1 electron (p^3 ⍰⍰⍰).
- Extra electron-electron repulsions make it easier to lose electron from p^4 than p^3.

ELECTRONEGATIVITY

								H 2.1									He
Li 1.0	Be 1.5											B 2.0	C 2.5	N 3.0	O 3.5	F 4.0	Ne
Na 0.9	Mg 1.2											Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0	Ar
K 0.8	Ca 1.0	Sc 1.3	Ti 1.5	V 1.6	Cr 1.6	Mn 1.5	Fe 1.8	Co 1.8	Ni 1.8	Cu 1.9	Zn 1.6	Ga 1.6	Ge 1.8	As 2.0	Se 2.4	Br 2.8	Kr
Rb 0.8	Sr 1.0	Y 1.2	Zr 1.4	Nb 1.6	Mo 1.8	Tc 1.9	Ru 2.2	Rh 2.2	Pd 2.2	Ag 1.9	Cd 1.7	In 1.7	Sn 1.8	Sb 1.9	Te 2.1	I 2.5	Xe
Cs 0.7	Ba 0.9	La 1.1	Hf 1.3	Ta 1.5	W 1.7	Re 1.9	Os 2.2	Ir 2.2	Pt 2.2	Au 2.4	Hg 1.9	Tl 1.8	Pb 1.8	Bi 1.9	Po 2.0	At 2.2	Rn

Power of an atom to attract the 2 electrons in a covalent bond

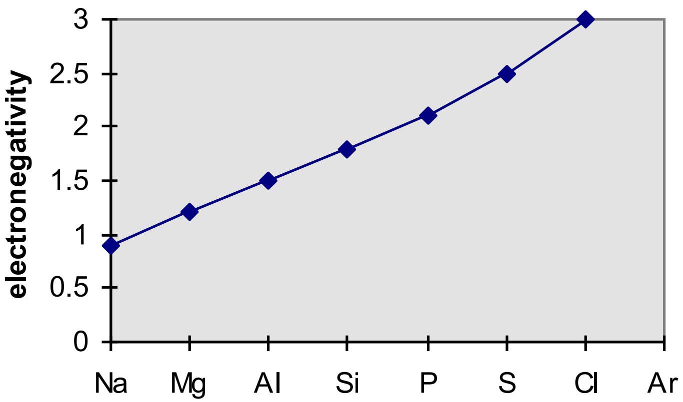
ELECTRONEGATIVITY



Power of an atom to attract the 2 electrons in a covalent bond

ELECTRONEGATIVITY

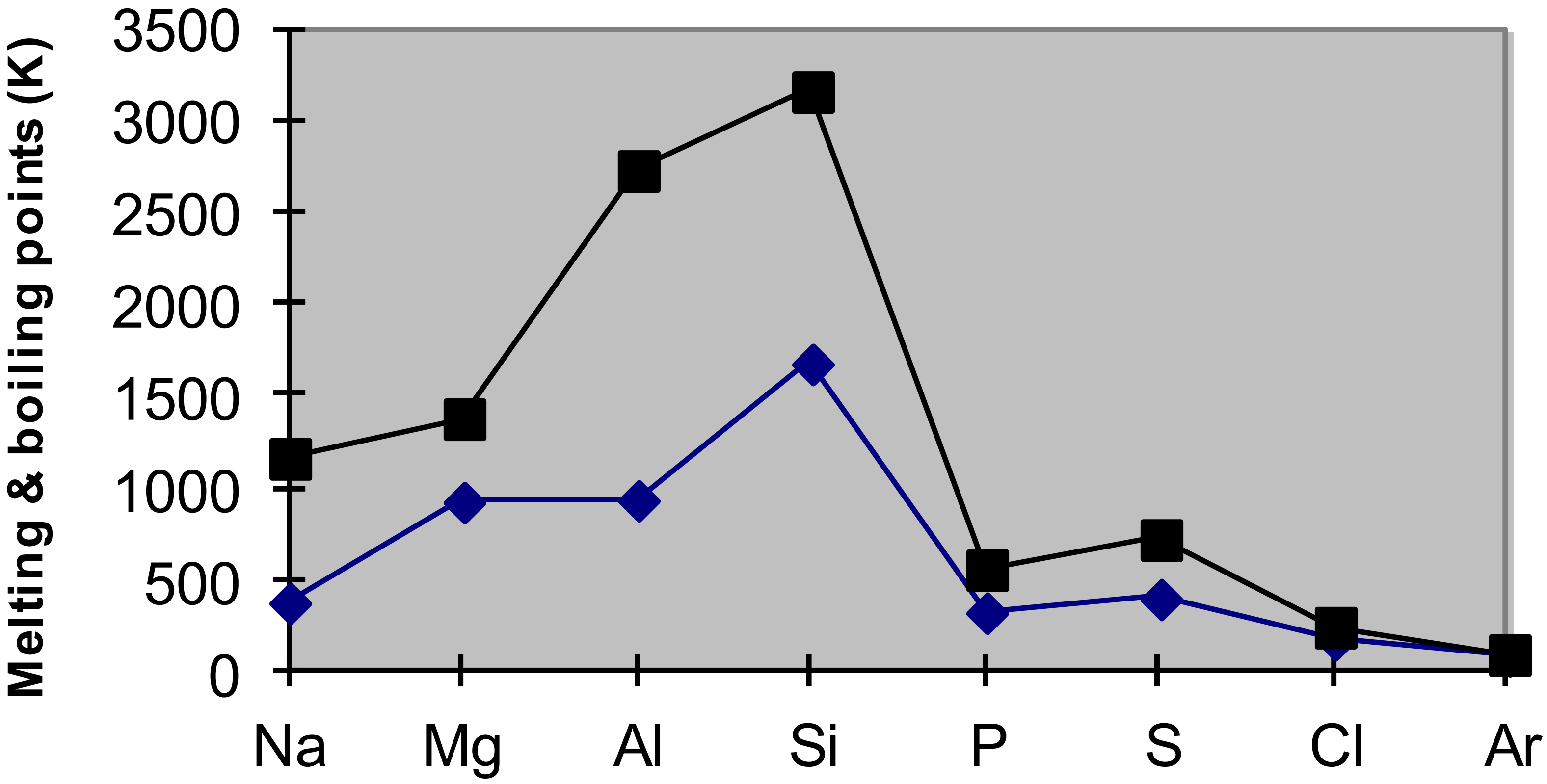
Power of an atom to attract the 2 electrons in a covalent bond



Across a period:

- more protons in nucleus
- smaller atomic radius
- so stronger attraction between nucleus and 2 electrons in covalent bond

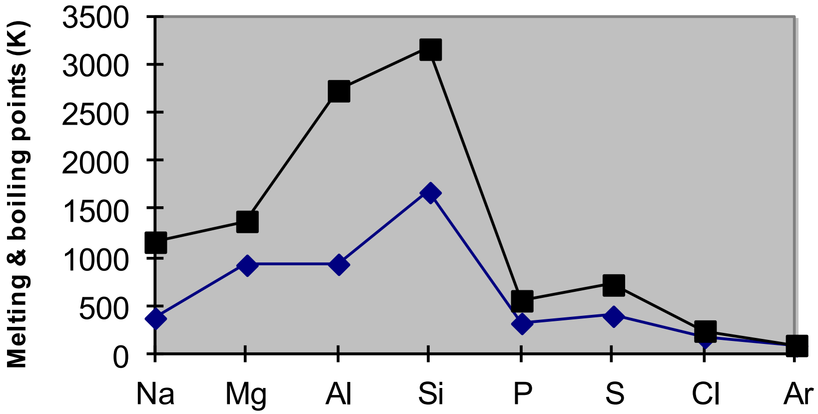
MELTING & BOILING POINTS

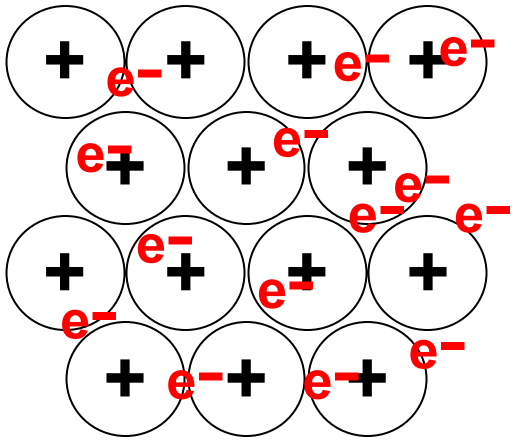


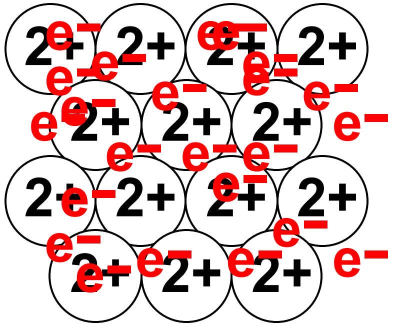
MELTING & BOILING POINTS

Na	**metallic**
Mg	**metallic**
Al	**metallic**
Si	**giant covalent**
P_4	**simple molecular**
S_8	**simple molecular**
Cl_2	**simple molecular**
Ar	**monatomic**

- Strong attraction between metal ions and delocalised electrons
- Al > Mg > Na as
  - ✓ higher charge
  - ✓ more delocalised electrons
  - ✓ smaller ions







Sodium (Na)
Mpt 98ºC

Magnesium (Mg)
Mpt 650ºC

MELTING & BOILING POINTS

Na	**metallic**	• Strong attraction between metal ions and delocalised electrons • Al > Mg > Na as ✓ higher charge ✓ more delocalised electrons ✓ smaller ions
Mg	**metallic**	
Al	**metallic**	
Si	**giant covalent**	• Have to break many strong covalent bonds
P_4	**simple molecular**	• Weak van der Waals' forces between molecules • $S_8 > P_4 > Cl_2$ (bigger molecules, more electrons, more vdW)
S_8	**simple molecular**	
Cl_2	**simple molecular**	
Ar	**monatomic**	• Very weak van der Waals' forces between atoms