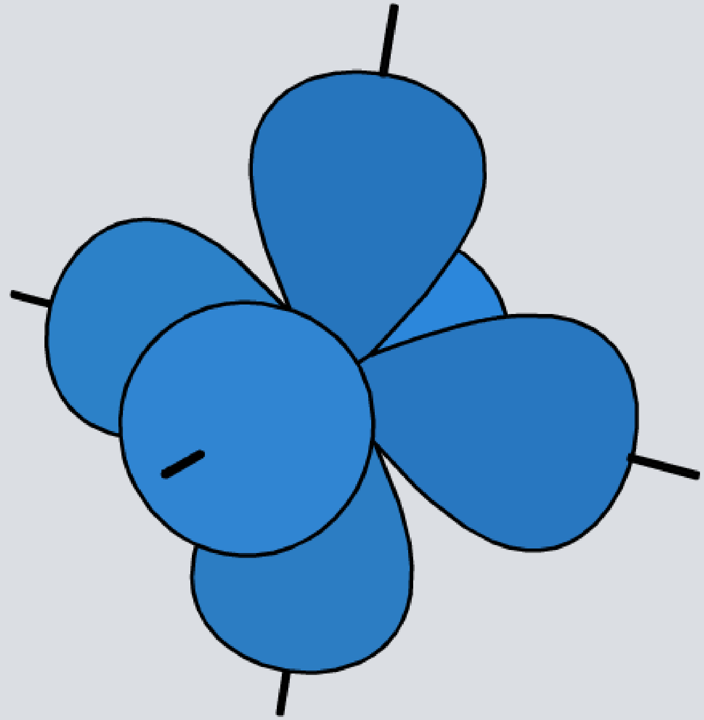
Orbitals



Electron orbitals

It is impossible to exactly locate the position of an electron within an energy sublevel. By measuring the electron density around the nucleus, it is possible to define regions where electrons are most likely to be found at any one time. These regions are called **orbitals**.

Each energy sublevel has one or more orbitals, each of which can contain a maximum of two electrons.

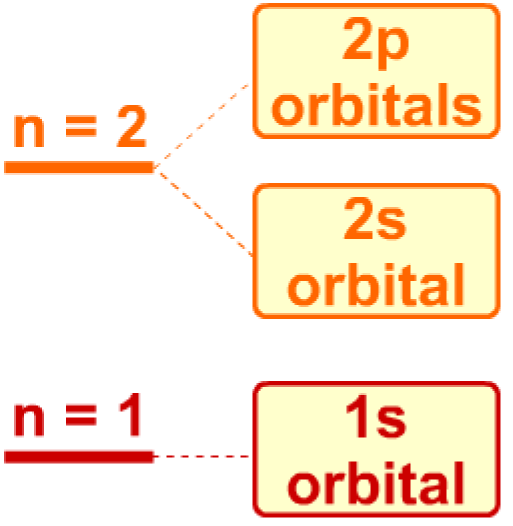
sublevel	no. orbitals	max. no. electrons
s	1	2
p	3	6
d	5	10
f	7	14

What are the shapes of electron orbitals?

Electron orbitals are regions around the nucleus of an atom where there is a high probability of finding an electron.

All orbitals can hold a maxiumum of two electrons, but those in different sublevels are different shapes and sizes.

Click a button on the right to see the shapes of some **s** and **p** orbitals.

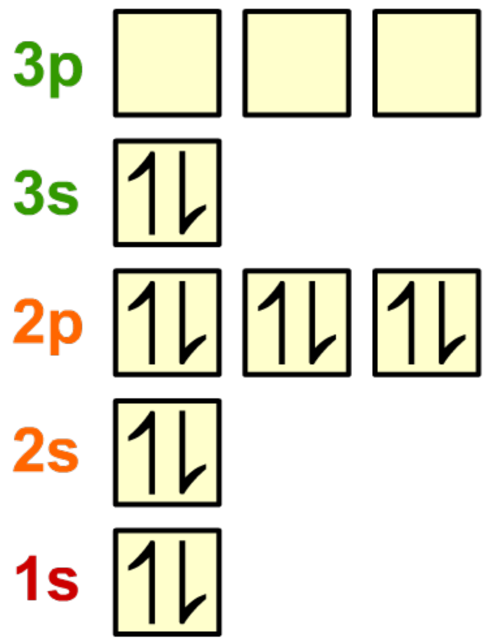


The Pauli exclusion principle and spin

The Pauli exclusion principle states that each orbital may contain no more than two electrons.

It also introduces a property of electrons called **spin**, which has two states: “up” and “down”. The spins of electrons in the same orbital must be opposite, i.e. one “up” and one “down”.

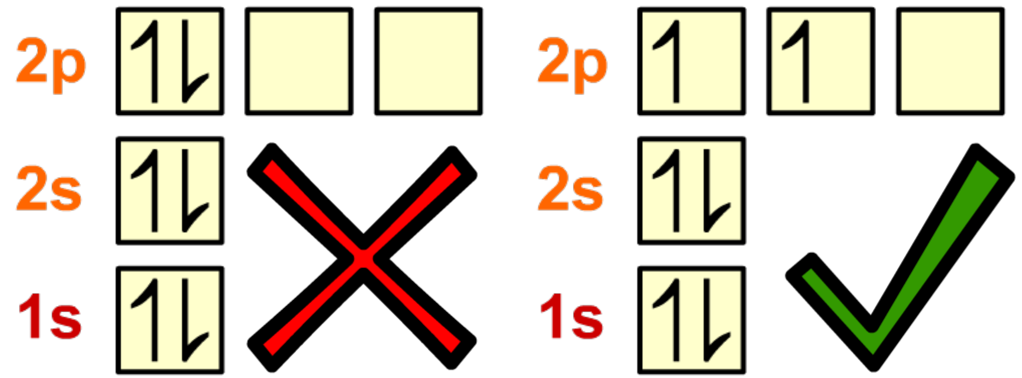
A **spin diagram** shows how the orbitals are filled. Orbitals are represented by squares, and electrons by arrows pointing up or down.



spin diagram for magnesium, $1s^2 2s^2 2p^6 3s^2$

Rules for filling electrons

When two electrons occupy a p sublevel, they could either completely fill the same p orbital or half fill two different p orbitals.

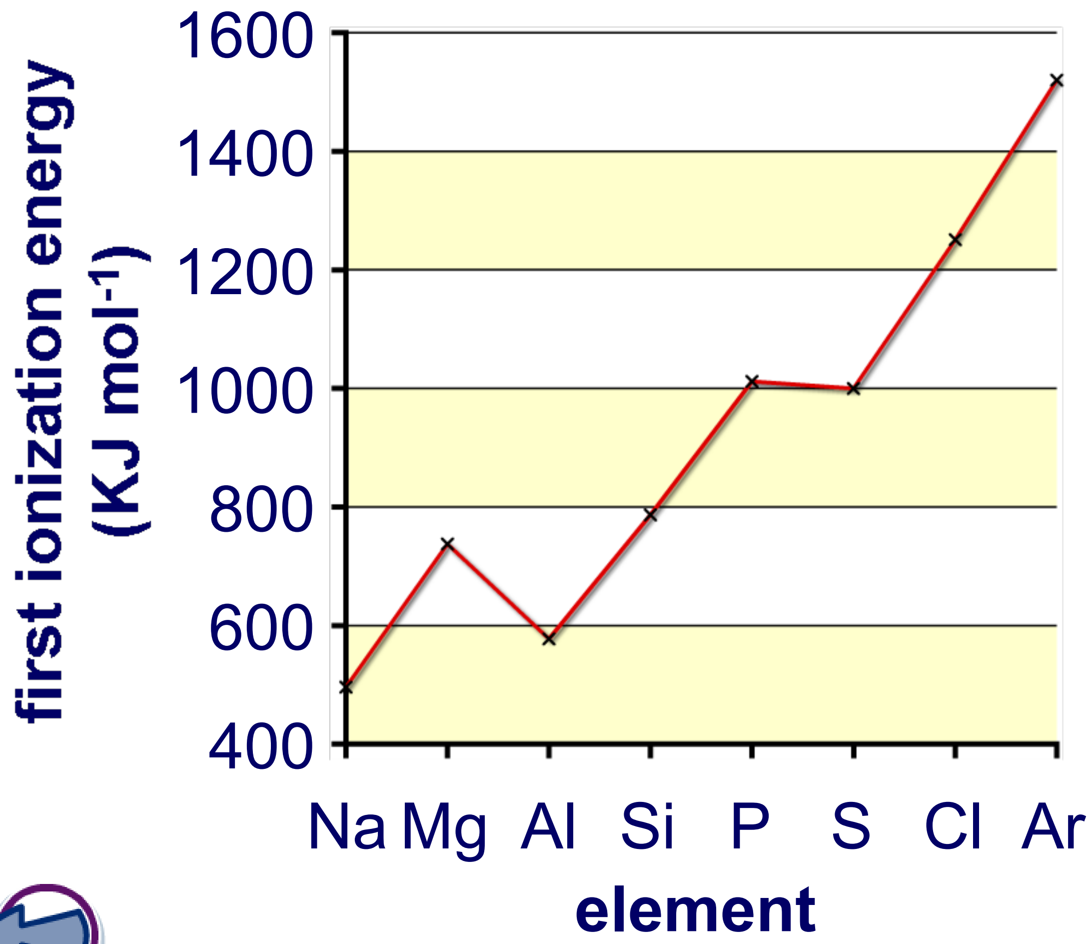


Hund's rule states that single electrons occupy all empty orbitals within a sublevel before they start to form pairs in orbitals.

If two electrons enter the same orbital there is repulsion between them due to their negative charges. The most stable configuration is with single electrons in different orbitals.

Evidence for Hund's rule

The first ionization energies for the elements in period 3 show a general increase.

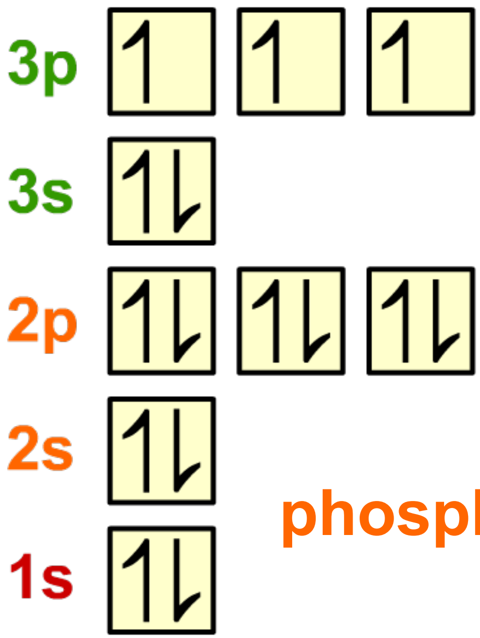


However, sulfur's value is below that of phosphorus. As the highest energy electrons of both are in the 3p sublevel this is evidence for **Hund's rule**.

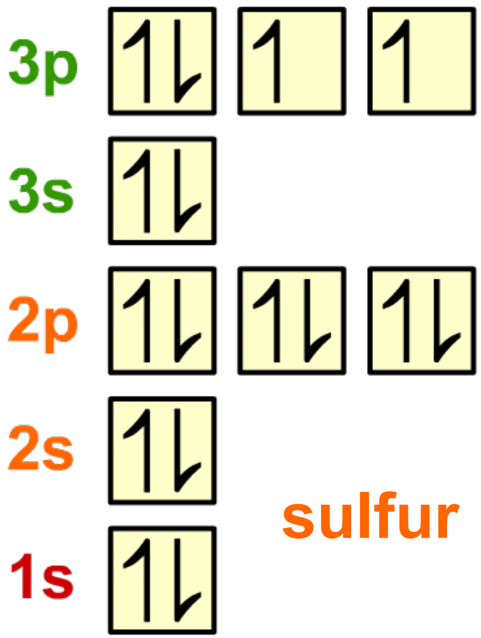
Evidence for Hund's rule: P vs. S

Phosphorus has three electrons in its 3p sublevel, and sulfur has four.

The lower first ionization energy for sulfur is because it has a pair of electrons in one of the 3p orbitals. Mutual repulsion between these two electrons makes it easier to remove one of them.



phosphorus



sulfur

Electron configuration of Cr and Cu

The electron configurations of chromium and copper are exceptions to the normal rules of orbital filling:

chromium	copper
$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^{10}$

In each case the 4s orbital contains one electron. This is because the 4s and 3d sublevels lie very close together in energy, and the 3d being either half full or completely full is a lower energy arrangement.

With larger atoms like this it can be useful to shorten the electron arrangement. Copper can be shortened to **$[Ar]4s^1 3d^{10}$**.

Creating electron spin diagrams

Click an element to start creating its spin diagram.

1.0 H 1																	4.0 He 2
6.9 Li 3	9.0 Be 4											10.8 B 5	12.0 C 6	14.0 N 7	16.0 O 8	19.0 F 9	20.2 Ne 10
23.0 Na 11	24.3 Mg 12											27.0 Al 13	28.1 Si 14	31.0 P 15	32.1 S 16	35.5 Cl 17	39.9 Ar 18
39.1 K 19	40.1 Ca 20	45.0 Sc 21	47.9 Ti 22	50.9 V 23	52.0 Cr 24	54.9 Mn 25	55.8 Fe 26	58.9 Co 27	58.7 Ni 28	63.5 Cu 29	65.4 Zn 30	69.7 Ga 31	72.6 Ge 32	74.9 As 33	79.0 Se 34	79.9 Br 35	83.8 Kr 36