

## Electron configurations of transition metal elements

### Hydrogen

$Z = 1$ . Its electron configuration is  $1s^1$ . Its electron diagram is on the right.



### Helium

$Z = 2$ . Its electron configuration is  $1s^2$ . Its electron diagram is on the right.



The  $1s$  sub-level is full, so completing the first principal energy level.

The  $n = 2$  level is used next.

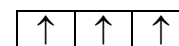
### Lithium

$Z = 3$ . Its electron configuration is  $1s^2 2s^1$ . Its electron diagram is on the right.



### Nitrogen

$Z = 7$ . Its electron configuration is  $1s^2 2s^2 2p^3$ . Its electron diagram is on the right.



Nitrogen obeys **Hund's multiplicity rule**, i.e. the favoured configuration is the one in which the electrons occupy different orbitals and have the same spins.



### Oxygen

$Z = 8$ . Its electron configuration is  $1s^2 2s^2 2p^4$ . Its electron diagram is on the right.

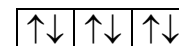


The fourth  $2p$  electron pairs with one of the other three  $2p$  electrons.



### Neon

$Z = 10$ . Its electron configuration is  $1s^2 2s^2 2p^6$  (see electron diagram on the right).

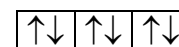


The  $2p$  sub-level is full, completing the  $n = 2$  level. The  $n = 3$  level is used next.



### Argon

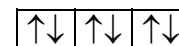
$Z = 18$ . Its electron configuration is  $1s^2 2s^2 2p^6 3s^2 3p^6$  (see diagram on the right).



This electron configuration can also be written as  $[\text{Ne}] 3s^2 3p^6$ , which saves space.



The  $3p$  sub-level is full, completing the  $n = 3$  level. The  $n = 4$  level is used next.



Potassium ( $Z = 19$ ) is  $[\text{Ar}] 4s^1$ , and Ca ( $Z = 20$ ) is  $[\text{Ar}] 4s^2$ . Once the  $4s$  sub-level is full, the next ten elements (the **d block** elements, Sc to Zn) use the  $3d$  orbitals.



### Transition metals

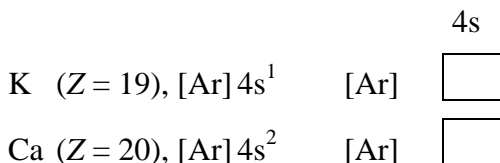
The d block elements contain the **transition metals**. These are elements which form some compounds in which there is an incomplete sub-level of d electrons. This means that strictly speaking scandium and zinc ( $[\text{Ar}] 3d^0$  and  $[\text{Ar}] 3d^{10}$  in compounds), and copper to some extent, are **not** transition metals. However, they are often included as their compounds resemble those of transition metals.

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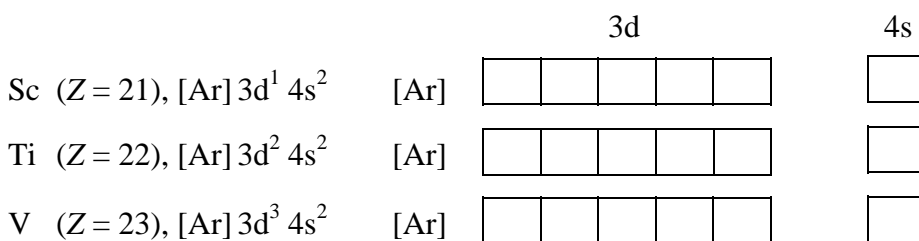
**Remember:**  $[\text{Ar}] = 1s^2 2s^2 2p^6 3s^2 3p^6$

**Note:** The orbitals of  $n = 4$  overlap with those of  $n = 3$ , i.e. in energy 4s below 3d, so after argon the 4s sub-level is filled first.

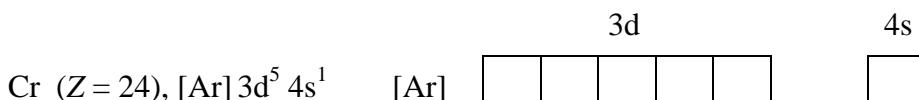
1. Complete the electron diagrams for potassium and calcium:



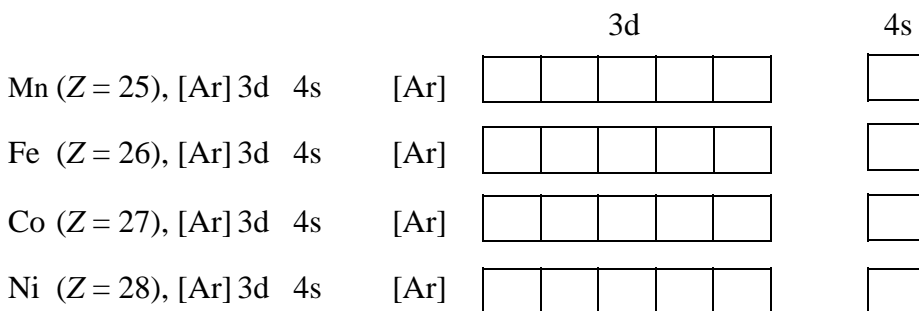
2. Now complete the electron diagrams for scandium, titanium and vanadium:



3. As far as vanadium, everything is straightforward. However, the electron configuration of the next element (chromium) is not  $[\text{Ar}] 3d^4 4s^2$  as you might expect. A half-filled d sub-level is more energetically favourable than a half-filled s sub-level, so one of the 4s electrons is promoted to a 3d orbital. Now complete the electron diagram for chromium:



4. For the next four elements, the 3d sub-level continues to fill. Complete the electron configurations and diagrams for manganese, iron, cobalt and nickel:



5. When we get to copper, the situation is similar to that found with chromium. In this case, a full 3d sub-level is more energetically favourable than a full 4s sub-level. Complete the electron configurations and diagrams for copper and zinc:

