Answer on Question #53306 – Chemistry – Inorganic Chemistry

Question

What is oxidation state?
How can find out the oxidation state of particular element?
Explain its trend in the group and period, give reasons

Answer

The oxidation state is an indicator of the degree of oxidation (loss of electrons) of an atom in a chemical compound. The oxidation state, which may be positive, negative or equal to zero, is the hypothetical charge that an atom would have if all bonds to atoms of different elements were completely ionic, with no covalent component.

To find out the oxidation state of particular element one should use some simple rules:

1. The oxidation state of an element in a simple substance (for example, He or Cl_2, or Fe, or C, or whatever containing one type of atoms) is equal to zero.

2. The sum of the oxidation states of all the atoms or ions in a neutral compound is zero.

3. The sum of the oxidation states of all the atoms in an ion is equal to the charge on the ion.

4. The more electronegative element in a substance is given a negative oxidation state. The less electronegative one is given a positive oxidation state.

5. Some elements almost always have the same oxidation states in their compounds:

Element	Oxidation state
Group 1 metals (Li, Na, K, Rb, Cs, Fr)	always +1
Group 2 metals (Be, Mg, Ca, Sr, Ba, Ra)	always +2
Fluorine (F)	always -1
Oxygen (O)	usually -2 (except in peroxides (-1) and F_2O (+2))
Hydrogen (H)	usually +1 (except in metal hydrides (-1))

Having known the oxidation states of these elements in the compound and having known the rule 3, the oxidation state of particular element can be found.

For example, let's find the oxidation state of Fe in Fe_2O_3. We know that oxidation state for O is -2 and the sum of the oxidation states equals 0. So, we have

$-2 \cdot 3 + Fe^{OS} \cdot 2 = 0$, whence $Fe^{OS} = 2 \cdot 3 / 2 = 3$ i.e. +3

The oxidation state trends in group and period.

Group number in periodic table shows number of electrons on the element's outermost subshell while period number shows number of subshells in the element.

The oxidation state of an element is related to the element's electronegativity, tendency of an atom to attract electrons towards it. The oxidation state tends to be positive in case of low electronegativity and negative in case of high electronegativity. Electronegativity increases in a period (from left to right) and decreases in a group (from top to bottom).

Despite the decrease in electronegativity in a group all the elements in it have the same characteristic oxidation states. So, the main point determining the oxidation state is electron configuration of an element, whose subshells may be filled to different extent.

Most atoms do not have completely filled subshells and will gain, lose, or share electrons with other atoms to "complete" these subshells. For example, the metals in Groups 1 and 2 readily form positively charged ions resulting from the loss of electrons as shown below

$$Li(1s^2\ 2s^1) \rightarrow Li^+(2s^1) + e^-$$

$$Mg(1s^2\ 2s^2\ 2p^6\ 3s^2) \rightarrow Mg^{2+}(1s^2\ 2s^2\ 2p^6) + 2\ e^-$$

The nonmetals in Groups 13-17 form negatively charged ions (anions) resulting from the gaining electrons to completely fill the outermost *p*-subshell

$$O(1s^2\ 2s^2\ 2p^4) + 2\ e^- \rightarrow O^{2-}(1s^2\ 2s^2\ 2p^6)$$

$$F(1s^2\ 2s^2\ 2p^5) + e^- \rightarrow Br^-(1s^2\ 2s^2\ 2p^6)$$

Other nonmetals can also acquire positive oxidation states due to the loss of the p- and the s-electrons. For example, in various compounds of sulfur, S ($1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$), it has a +4 oxidation state when having lost the $3p^4$ electrons and a +6 oxidation state when having lost the additional $3s^2$ electrons.

The oxidation state of the transition and inner transition elements is positive (they are metals with low electronegativity), but because some of the subshells for transition and inner transition metals are close in energy, there are often several different possible oxidation states.

Vanadium for example has electron configuration

$$V\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3\ 4s^2$$

written in order by increasing subshell number (period number in periodic table) which represents the order of removing electrons.

The removal of first electrons (s-electrons of the outermost subshell like in case of Group 2 metals) results in

$$V^{2+}\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^3$$

Removal of the electrons from partially filled subshell results in

$$V^{5+}\ 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$$

So, vanadium in its compounds normally has the oxidation state +2 or +5.

The oxidation state of other transition metals may be predicted the same way.