Oxidation States
Losers and Gainers and Relative Reactivity

Oxidation states indicate the "charge" for a particular atom. This "charge" could be the number of electrons that the atom has lost or gained in ionic bonding. Or the loss or gain may be partial, as is the case in covalent bonding (formal charge). However in either case the oxidation states will be critically useful in determining how atoms fit together to form compounds or polyatomic ions. Common oxidation states are listed on the periodic table. Examples are:

2+	2+	2–	1+
Mg	Zn	O	H

STOP AND FIND THESE NUMBERS ON THE PERIODIC TABLE – USE THE LEGEND

Some elements have multiple oxidation states (called polyvalent elements) examples from your periodic table and what is means are as follows:

±2,4,6		2–	2+	4+	6+	
S	➜	S	S	S	S	
±1,3,5,7		1–	1+	3+	5+	7+
Cl	➜	Cl	Cl	Cl	Cl	Cl
2,3,6		2+	3+	6+		
Fe	➜	Fe	Fe	Fe		

When multiple oxidation states are present, **there will be only one negative state and it will always follow the octet rule (see S and Cl above).**

Also all elements have the additional oxidation state of 0 (which is not listed on the periodic table because of its ubiquitous nature). The oxidation state of zero will only apply when the element is in pure elemental form:

	^{0}	^{0}	^{0}	^{0}
eg	O_2	N_2	Fe	Na

Oxidation States in Compounds and Ions:

For a compound to be neutral, the sum of its oxidation states must be zero

eg	$^{1+2-}H_2O$	2(1+)	+	(2−)	=	0
	$^{3-1+}NH_3$	(3−)	+	3(1+)	=	0
	$^{4+2-}CO_2$	(4+)	+	2(2−)	=	0
	$^{4-1+}CH_4$	(4−)	+	4(1+)	=	0

For an ion, the sum of the oxidation states must equal the net charge on the ion. This is simple for monatomic ions:

eg	$^{1+}Na^{1+}$	$^{2+}Mg^{2+}$	$^{3+}Al^{3+}$	$^{1+}H^{1+}$

NOTE THE DIFFERENCE BETWEEN THE CHARGE, WRITTEN TO THE UPPER RIGHT AND THE OXIDATION STATE, WRITTEN OVER THE ELEMENT SYMBOL

And more complex for polyatomic ions (we will meet lots of these later)

eg $^{6+2-}SO_4^{2-}$ (6+) + 4(2−) = 2−

$^{3-1+}NH_4^{1+}$ (3−) + 4(1+) = 1+

$^{5+2-}PO_4^{3-}$ (5+) + 4(2−) = 3−

$^{2-1+}OH^{1-}$ (2−) + (1+) = 1−

Further Examples to Make Chemical Formula and Balanced Chemical Equations:

$$^{0}2Cs + {}^{1+\ 2-}2H_2O \rightarrow {}^{1+\ \ 2-1+}2Cs^{1+}OH^{1-} + {}^{0}H_2$$

$$^{0}Ba + {}^{1+2-}2H_2O \rightarrow {}^{2+\ \ 2-1+}Ba^{2+}(OH)^{1-}_{2} + {}^{0}H_2$$

$$^{0}2La + {}^{1+2-}6H_2O \rightarrow {}^{3+\ \ 2-1+}2La^{3+}(OH)^{1-}_{3} + {}^{0}3H_2$$

NOTE THE USE OF IONIC CHARGES IN THE THIRD COMPOUND IN EACH EQUATION – STUDY THIS – THE OXYGENS (2−) AND THE HYDROGENS (1+) ADD TO GIVE THE HYDROXIDES (1−) [HYDROXIDE IS AN OH^{1-} ION]

Losers, Gainers and Oxidation States:

- losers will have an increase in oxidation state (positive direction)
- gainers will have a decrease in oxidation state (negative direction)

$$\overset{0}{2Na} + \overset{0}{Cl_2} \rightarrow \overset{1+\ 1-}{2Na^{1+}Cl^{1-}}$$

- each sodium has lost $1e^{1-}$
- each chlorine has gained $1e^{1-}$

Losers vs Gainers:

Almost all chemical reactions can be thought of as loser/gainer reactions. The element with the greatest tendency to gain electrons (which will have the higher electronegativity) will pull electrons from the element with the lower ionization energy (ionization energy and electronegativity are more or less parallel). Comparison of electronegativity is usually the best way to figure out who loses and who gains. The losing and gaining will follow the octet rule most of the time.

Lose
Electrons
Oxidized

the
lion
says

Gain
Electrons
Reduced

Relative Reactivity:

Relative reactivity can best be explained by the following steps:

- determine if you are comparing the relative reactivity of losers or gainers
- use the correct relevant atomic property
  - for losers ➜ consider ionization energy
  - for gainers ➜ consider electronegativity
- then use a comparison of the relevant atomic property
  - for losers ➜ lower ionization energy means easier losing and therefore faster reaction
  - for gainers ➜ higher electronegativity means better gaining abilities and therefore faster reaction