

Redox Reactions

Reduction-Oxidation Reaction — reaction involving transfer of electrons.

reductions means GAINS electrons

oxidation — LOSES electrons

OILRIG

Oxidation #s (aka Oxidation State)

* The # of electrons an element has gained or lost (typically...)

Neg oxidation #'s → gained electrons
charge more neg
* reduced

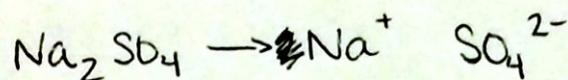
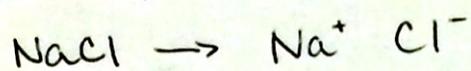
Pos oxidation #'s → lost electrons
charge more pos
* oxidized

worth Noting → Fluorine is most electronegative element.

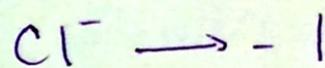
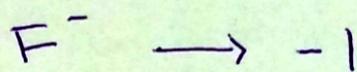
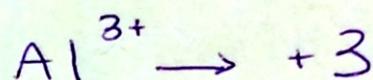
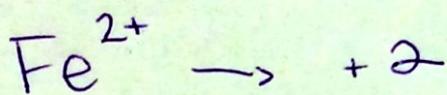
Oxygen is 2nd

Rules for Finding Oxidation States / #s

* in general, think of ionic compounds as their separate ions. → metals + non-metals



① For monatomic ion - oxidation is equal to charge (magnitude & direction)



② Hydrogen is +1 EXCEPT in metal halides (a metal + hydrogen NaH)

③ Oxygen is -2 EXCEPT in Peroxides (H_2O_2), it will be -1 AND with Fluorine alone, (it be +)

④ In elemental form (Li , Cl_2) even if diatomic the oxidation # is \emptyset .

⑤ For a neutral compound, the sum of oxidation #'s is \emptyset .

⑥ For a polyatomic ion (PO_4^{3-}) the sum of oxidation # is equal to charge.

* Oxidation # - don't count subscripts

only use subscripts to find sum of ox #'s & balance compounds

Hints + Different Approaches ...

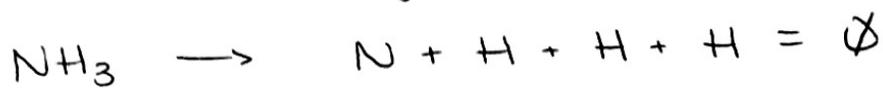
Oxidation States

* remember the oxidation # is telling us about gain/loss of electrons per atom.

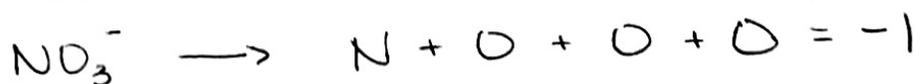
That's why subscripts aren't factored in when finding oxidation #s or their changes.

- we can use sum of oxidation #s to solve for unknown #s (now we need those subscripts)

→ can think of it as solving for unknown (see neutral compounds to 0, polyatomic ions to their charge)



or



* You'll need the naming compounds + common ions handout for this unit 3 test.