

Announcements for Thursday, 07NOV2024

- none

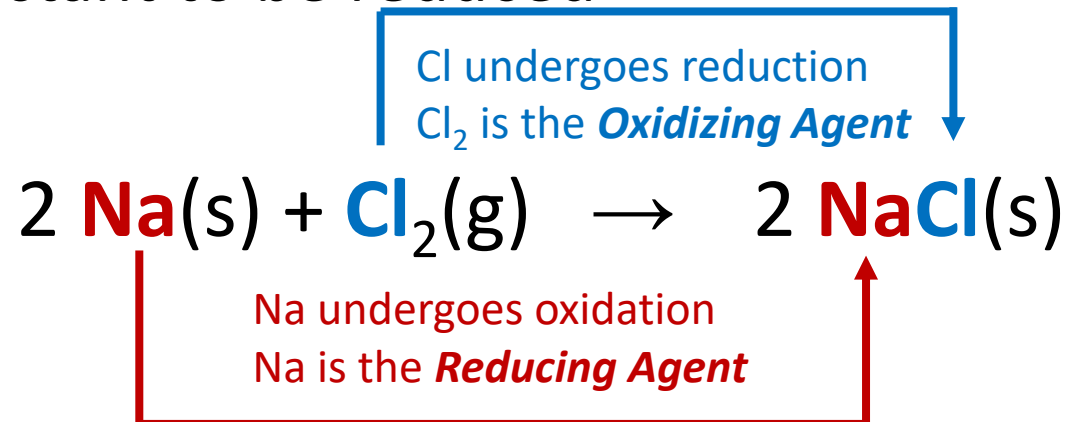
ANY GENERAL QUESTIONS? Feel free to see me after class!

Try These On Your Own

- 10.0 g $\text{Ca(OH)}_2(\text{s})$ was needed to completely neutralize 266 mL $\text{HCl}(\text{aq})$. What was the molarity of $\text{HCl}(\text{aq})$? **1.01 M**
- What volume of 2.5 M $\text{NaOH}(\text{aq})$ must be added to 100. mL of 0.50 M $\text{H}_3\text{PO}_4(\text{aq})$ to completely neutralize the acid? **60 mL**

Oxidation-*Red*uction Reactions (Redox Reactions)

- **oxidation-reduction reaction** = an electron-transfer reaction
 - electrons are transferred from one reactant to another
 - many important and relevant reactions are redox reactions
 - combustion, rusting/corrosion, batteries, metabolism, photosynthesis...
- **oxidation** = the loss of electrons; **reduction** = the gain of electrons
- **oxidizing agent** = the *reactant* that gets reduced during the process and allows another reactant to be oxidized
- **reducing agent** = the *reactant* that gets oxidized during the process and allows another reactant to be reduced



Oxidation-Reduction Reactions (Redox Reactions) (continued)

- redox reactions are comparatively harder to identify than precipitation reactions and acid-base reactions
- to correctly identify a redox reaction you must assign **oxidation numbers (?!?)** to all atoms and ions and look for changes as reactants go to products
 - the same species could be undergoing both oxidation and reduction at the same time
 - a *disproportionation reaction*



bromine is being both oxidized and reduced

Oxidation States/Numbers

- **oxidation number** = the *fictitious* charge an atom would have in a compound if all electrons were assigned to the more electronegative atom in that compound
 - oxidation numbers are assigned to help keep track of the transfer of electrons during a redox reaction
- for ions in an ionic compound, oxidation states have the same value as the ions' charges
 - NaCl
- for neutral atoms in a molecular compound, oxidation states are NOT ion charges
 - H₂O vs. OF₂
 - the more electronegative atom(s) gets full ownership of shared electrons
- oxidation numbers are IMAGINARY charges that are assigned based on rules
 - as opposed to ion charges which are real and measurable

Rules for Assigning Oxidation Numbers

- oxidation numbers are assigned to every atom or ion in a compound according to a specific procedure
 - oxidation numbers can be positive, negative, or fractional

the rules for assigning oxidation numbers MUST be applied in a specific order since certain rules take priority over others

1. the sum of oxidation numbers for each atom or ion in a compound must equal the overall charge of the compound
 - atoms in elemental forms have oxidation numbers of zero: $\text{H}_2(\text{g})$, $\text{O}_3(\text{g})$, $\text{Na}(\text{s})$, $\text{P}_4(\text{s})$, etc.
 - monatomic ions have oxidation numbers equal to their charge: Cu^+ ox # = +1, S^{2-} ox # = -2, etc.
2. Assign oxidation numbers of **+1** to **Group 1A metal cations** and **+2** to **Group 2A metal cations**
 - Li^+ , Na^+ , K^+ , Rb^+ , Cs^+ , Mg^{2+} , Ca^{2+} , Sr^{2+} , Ba^{2+}
3. Assign **fluorine** an oxidation number of **-1**

Rules for Assigning Oxidation Numbers (continued)

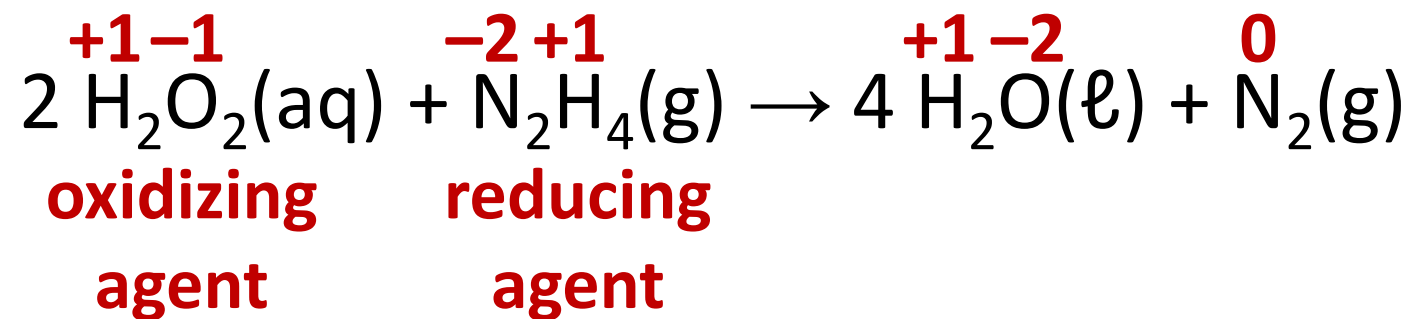
4. Assign **hydrogen** an oxidation number of **+1** *unless otherwise dictated by previous rules*
 - NaH, CaH₂
5. Assign **oxygen** an oxidation number of **-2** *unless otherwise dictated by previous rules*
 - H₂O₂, OF₂, NaO₂
6. Assign **Group 7A nonmetals (Cl, Br, I)** oxidation numbers of **-1** *unless otherwise dictated by previous rules*
7. Assign **Group 6A nonmetals (S, Se, Te)** oxidation numbers of **-2** *unless otherwise dictated by previous rules*
8. Assign **Group 5A nonmetals (P, As, Sb)** oxidation numbers of **-3** *unless otherwise dictated by previous rules*

The rules for assigning oxidation numbers will NOT be provided on an exam and must be MEMORIZED!

Minimum and Maximum Oxidation Numbers

- some elements can have many oxidation states
 - for **nitrogen**, minimum ox # = Group # – 8 ($5-8 = -3$); maximum ox # = Group # (**+5**)
 - for **sulfur**, minimum ox # = Group # – 8 ($6-8 = -2$); maximum ox # = Group # (**+6**)
 - for **chlorine**, minimum ox # = Group # – 8 ($7-8 = -1$); maximum ox # = Group # (**+7**)
- when a species has its maximum ox #, it can **only** be reduced (i.e., gain electrons) and act as an oxidizing agent
 - NO_3^- , SO_4^{2-} , ClO_4^-
- when a species has its minimum ox #, it can **only** be oxidized (i.e., lose electrons) and act as a reducing agent
 - NH_3 , H_2S , HCl
- when a species has an intermediate ox #, it can act as **either** an oxidizing agent (and be reduced) or as a reducing agent (and be oxidized)
 - NO , SO_2 , ClO^-
- transition metals can also have multiple ox #s, but can't be easily predicted from Group #

Identify the oxidizing agent and the reducing agent in the following reaction.

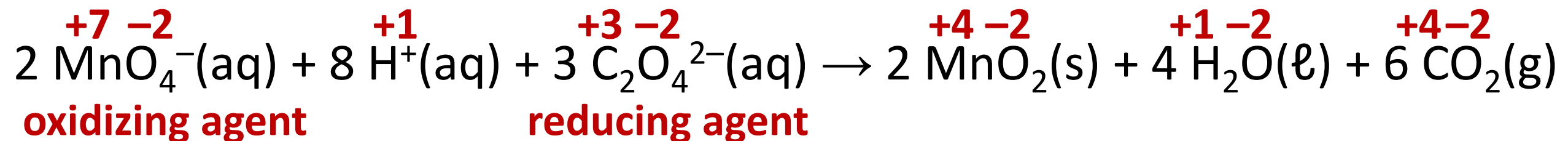


H: +1 \rightarrow +1 ... no change

O: -1 \rightarrow -2 ... gain of electrons ... reduction ... H_2O_2 is the oxidizing agent

N: -2 \rightarrow 0 ... loss of electrons ... oxidation ... N_2H_4 is the reducing agent

Identify the oxidizing agent and the reducing agent in the following reaction.



Mn: $+7 \rightarrow +4$... gain of electrons ... reduction ... MnO_4^- is the oxidizing agent

O: $-2 \rightarrow -2$... no change H: $+1 \rightarrow +1$... no change

C: $+3 \rightarrow +4$... loss of electrons ... oxidation ... $\text{C}_2\text{O}_4^{2-}$ is the reducing agent

Try These On Your Own

Assign oxidation numbers to all species and determine if the reaction is a redox. If it is, identify the oxidizing agent and the reducing agent.

