Announcements for Thursday, 07NOV2024

- none

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

Try These On Your Own

- 10.0 g $Ca(OH)_2$(s) was needed to completely neutralize 266 mL HCl(aq). What was the molarity of HCl(aq)? **1.01 M**

- What volume of 2.5 M NaOH (aq) must be added to 100. mL of 0.50 M H_3PO_4 (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

Cl undergoes reduction
Cl_2 is the ***Oxidizing Agent***

$$2\ Na(s) + Cl_2(g) \longrightarrow 2\ NaCl(s)$$

Na undergoes oxidation
Na is the ***Reducing Agent***

Oxidation-***Red***uction 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*

$$3\ BrF \longrightarrow BrF_3 + Br_2$$

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
  - H_2O vs. OF_2
    - 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: $H_2(g)$, $O_3(g)$, $Na(s)$, $P_4(s)$, 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, CaH_2
5. Assign oxygen an oxidation number of **–2** *unless otherwise dictated by previous rules*
   - H_2O_2, OF_2, NaO_2
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 <u>maximum ox #</u>, it can **only** be reduced (i.e., gain electrons) and act as an <u>oxidizing agent</u>
  - NO_3^-, SO_4^{2-}, ClO_4^-
- when a species has its <u>minimum ox #</u>, it can **only** be oxidized (i.e., lose electrons) and act as a <u>reducing agent</u>
  - NH_3, H_2S, HCl
- when a species has an <u>intermediate ox #</u>, 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.

$$2\,\overset{+1\,-1}{H_2O_2}(aq) + \overset{-2\,+1}{N_2H_4}(g) \longrightarrow 4\,\overset{+1\,-2}{H_2O}(\ell) + \overset{0}{N_2}(g)$$

oxidizing agent (H_2O_2) **reducing agent** (N_2H_4)

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.

$$2\,\underset{\textbf{oxidizing agent}}{\overset{+7\;\;-2}{MnO_4^-}}(aq) + 8\,\overset{+1}{H^+}(aq) + 3\,\underset{\textbf{reducing agent}}{\overset{+3\;-2}{C_2O_4^{2-}}}(aq) \longrightarrow 2\,\overset{+4\;-2}{MnO_2}(s) + 4\,\overset{+1\;-2}{H_2O}(\ell) + 6\,\overset{+4-2}{CO_2}(g)$$

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

O: $-2 \longrightarrow -2$ … no change   H: $+1 \longrightarrow +1$ … no change

C: $+3 \longrightarrow +4$ … loss of electrons … oxidation … $C_2O_4^{2-}$ 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.

$2\ NO(g) + 5\ H_2(g) \rightarrow 2\ NH_3(g) + 2\ H_2O(g)$

$N_2O_5\ (aq) + H_2O(\ell) \rightarrow 2\ HNO_3(aq)$

$5\ Fe^{2+}(aq) + 8\ H^{+}(aq) + MnO_4^{-}(aq) \rightarrow 5\ Fe^{3+}(aq) + Mn^{2+}(aq) + 4\ H_2O(\ell)$