

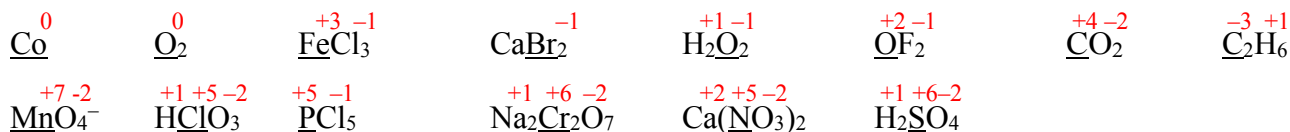
Notes 4-3: Ch. 4.3-4.4: Acid-Base Reactions; Oxidation Number, Half-Reactions

- Arrhenius Acids & Bases
 - Arrhenius Acids produce H^+ in aqueous solutions: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$
 - Arrhenius Bases produce OH^- in aqueous solutions: $NaOH(aq) \rightarrow Na^+(aq) + OH^-(aq)$
- Brønsted Acids & Bases
 - Brønsted Acids *donate* protons (H^+): $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$
 - H_3O^+ is *hydronium* ion, better representation of species than H^+
 - Brønsted Bases *accept* protons: $NH_3(aq) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$
 - H_2O is *amphoteric*—it behaves as both an acid and a base
 - Any $R_3N + H^+ \rightarrow R_3NH^+$
 - Acids classified by number of protons:
 - Monoprotic has one proton:
 - Strong monoprotic acids: $HCl, HBr, HI, HClO_4, HNO_3$
 - All others weak: HF, HNO_2, CH_3COOH (any acid ending in $COOH$), etc.
 - Polyprotic acids: Diprotic has two protons: $H_2SO_4, H_2C_2O_4$
 - H_2SO_4 is strong, but only for first dissociation:
 - Polyprotic acids dissociate one at a time in H_2O :
 - $H_2SO_4 + H_2O \rightarrow HSO_4^- + H_3O^+$ • Intermediate species (HSO_4^-) amphoteric
 - $HSO_4^- + H_2O \rightarrow SO_4^{2-} + H_3O^+$
 - All others weak: H_2SO_3, H_2CO_3 , etc.
 - Triprotic has three protons: H_3PO_4 (none strong)
 - Both intermediate species ($H_2PO_4^-, HPO_4^{2-}$) amphoteric
 - Strong bases: Hydroxides of alkali metals (e.g. $NaOH$), $Ca(OH)_2$, $Sr(OH)_2$, & $Ba(OH)_2$ (same as exceptions to OH^- solubility rules). **Cations for “b” on periodic table**
- Neutralization Reactions: Double Replacement w/acid & base (one usually strong)
 - acid + base \rightarrow **water + salt** (any ionic compound not containing H^+ or OH^-)
 - $HBr(aq) + KOH(aq) \rightarrow KBr(aq) + H_2O(l)$
 - Ionic equation: $H^+(aq) + Br^-(aq) + K^+(aq) + OH^-(aq) \rightarrow K^+(aq) + Br^-(aq) + H_2O(l)$
 - Net ionic equation: $H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$
 - Be careful—weak acids & bases do NOT dissociate:
 - $HF(aq) + NaOH(aq) \rightleftharpoons NaF(aq) + H_2O(l)$
 - Ionic equation: $HF(aq) + Na^+(aq) + OH^-(aq) \rightleftharpoons Na^+(aq) + F^-(aq) + H_2O(l)$
 - Net ionic equation: $HF(aq) + OH^-(aq) \rightleftharpoons F^-(aq) + H_2O(l)$
 - $H_2SO_3(aq) + 2 LiOH(aq) \rightarrow 2 H_2O(l) + Li_2SO_3(aq)$
 - Ionic equation: $H_2SO_3(aq) + 2 Li^+(aq) + 2 OH^-(aq) \rightarrow 2 H_2O(l) + 2 Li^+(aq) + SO_3^{2-}(aq)$
 - Net ionic equation: $H_2SO_3(aq) + 2 OH^-(aq) \rightarrow 2 H_2O(l) + SO_3^{2-}(aq)$
- Oxidation-Reduction (Redox) Reactions
 - Many reactions involve transfer of electrons:
 - One species loses electrons, it is *oxidized* (**L**osing **E**lectrons is **O**xidation-LEO)
 - One species gains electrons, it is *reduced* (**G**aining **E**lectrons is **R**eductions-GER)
 - Mnemonic: **LEO the Lion says GER**
 - Ox-Red occur simultaneously and MUST occur together—can’t have one without the other.
- Oxidation Number
 - Bookkeeping location of electrons: Number of charges an atom would have *if electrons were transferred completely*.
 - Rules: since sharing electrons is involved, electronegativity is important.
 - Write ox. number for *single* atom

Rules for Assigning Oxidation Numbers

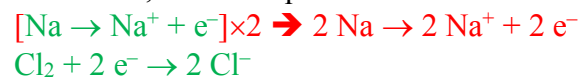
- All atoms in their elemental state (H_2 gas, Li metal, solid C) have an oxidation number of 0.
- All monatomic ions (whether aqueous or in a solid ionic compound) have an oxidation number equal to their charge (e.g. K^+ , Br^- , Fe^{3+} , S^{2-}).
- All neutral compounds have a net oxidation number of 0. All polyatomic ions have a net oxidation number equal to their charge (e.g. in NO_3^- , the sum of the oxidation numbers of N and the 3 O atoms must equal -1).
- Fluorine is the most electronegative element, so in compounds its oxidation number is *always* -1 .
- Oxygen's oxidation number is *usually* -2 , except:
 - When combined with F (which is more electronegative), F is -1 and O follows rule 3.
 - In peroxides (O_2^{2-}) its oxidation number is -1 (e.g. Na_2O_2)
 - In superoxides (O_2^-) its oxidation number is $-\frac{1}{2}$ (rare) (e.g. KO_2)
- Hydrogen has oxidation number $+1$ in all compounds except metal hydrides, such as LiH, and specialty compounds such as sodium borohydride, NaBH_4 , where its oxidation number is -1 since it is more electronegative.
- The oxidation number of all other atoms in a species must be determined according to the above rules. For example, in NO_3^- , the oxidation number of each O is -2 and the sum of the oxidation numbers must be -1 , so the oxidation number of N must be $+5$).
 - In any binary compound, the more electronegative atom is assigned its normal negative charge. E.g. in SCl_6 , Cl is assigned -1 , so S is $+6$ since the molecule is neutral.

Practice: Find the oxidation number of the underlined atom in the following species:



- Half Reactions

- Redox reactions can be separated into oxidation and reduction half-reactions
- Half reactions show electron transfer explicitly.
- e.g. $2 \text{Na} + \text{Cl}_2 \rightarrow 2 \text{NaCl}$
 - Assign oxidation numbers: $2 \text{Na}^0 + \text{Cl}_2^0 \rightarrow 2 \text{Na}^+ \text{Cl}^-$
 - Split into ox & red $\frac{1}{2}$ -reactions, balance mass, then balance charge using electrons
 - Oxidation Reaction: Na *is oxidized*: $\text{Na} \rightarrow \text{Na}^+ + \text{e}^-$ Note: charge must balance but does not need to be 0
 - Reduction Reaction: Cl_2 *is reduced*: $\text{Cl}_2 + 2 \text{e}^- \rightarrow 2 \text{Cl}^-$
 - Na donates electrons to Cl_2 as it is *oxidized*—Na reduces Cl_2 so it is *reducing agent*
 - Cl_2 accepts electrons from Na as it is *reduced*— Cl_2 oxidized Na so it is *oxidizing agent*
 - To recombine, need to equalize electrons:



Multiply by 2 to equalize & cancel electrons when

Cancel electrons and combine ions to get original equation.