

Competition for Electrons

Aim

- write equations for oxidation and reduction half reactions

Notes

Atoms compete for each other's electrons

- When chemical bonds form, electrons are either lost, gained or shared
- Oxidation-Reduction reactions (Redox reactions)
 - Metals
 - lose electrons (OXIDATION)[NOTE: as when metals combine with oxygen]
 - are oxidized
 - are reducing agents
 - Nonmetals
 - gain electrons reducing their oxidation states (REDUCTION)
 - are reduced
 - are oxidizing agents

Oxidation
Is
Loss

Reduction
Is
Gain

- Example 1 - $2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$

Mg	O ₂
★ loses electrons	★ gains electrons
★ gets oxidized to Mg ²⁺	★ gets reduced to O ²⁻
★ is the reducing agent for O ₂	★ is the oxidizing agent for Mg

- Half reactions — reaction showing either a gain or loss of electrons
 - $2\text{Mg}^0 \rightarrow 2\text{Mg}^{2+} + 4e^-$
 - $\text{O}_2^0 + 4e^- \rightarrow 2\text{O}^{2-}$
- Net equation (REDOX REACTION)— combination of the half reactions such that the number of electrons lost equals the number of electrons gained

$$2\text{Mg}(s) + \text{O}_2(g) \rightarrow 2\text{MgO}(s)$$
- Example 2 - More active metals replace less active metals in compounds by transferring electrons to them
 - Sample Reaction:

$$\text{Zn}(s) + \text{Cu}(\text{NO}_3)_2(aq) \rightarrow \text{Zn}(\text{NO}_3)_2(aq) + \text{Cu}(s)$$
 - Half reactions — reaction showing either a gain or loss of electrons
 - $\text{Zn}^0 \rightarrow \text{Zn}^{2+} + 2e^-$
 - $\text{Cu}^{2+} + 2e^- \rightarrow \text{Cu}^0$
 - Net equation — combination of the half reactions such that the number of electrons lost equals the number of electrons gained

$$\text{Cu}^{2+} + \text{Zn}^0 \rightarrow \text{Zn}^{2+} + \text{Cu}^0$$
 - Spectator ions — ions that are present during a reaction but do not participate in the reaction:

$$2\text{NO}_3^-$$

Oxidation number (Oxidation state) - number assigned to keep track of electrons based on the arbitrary assumption that shared electrons belong to the more electronegative element

- Rules for assigning oxidation numbers
 - Oxidation numbers for atoms that are free elements are always zero
 - The oxidation numbers of ions are the same as the charge on the ion
 - Some elements have only one oxidation state
 - group 1 metals always form 1+ ions and always have a +1 oxidation state
 - group 2 metals always form 2+ ions and always have a +2 oxidation state
 - Some elements usually have a particular oxidation state
 - oxygen has a -2 oxidation state except in peroxides where it is -1 and in compounds with fluorine (OF₂) where it is +2
 - hydrogen has a +1 oxidation state except in hydrides with group 1 and group 2 metals
 - the sum of the oxidation numbers
 - in a compound it is always zero
 - in a polyatomic ion it is equal to the charge on the ion
- Finding oxidation numbers
 - apply the rules
 - construct a table if necessary

Sample Problem

Find the oxidation state of the elements in K₂Cr₂O₇.

Element	K	Cr	O	TOTAL
Subscript	2	2	7	
Oxidation state	+1	?	-2	
Sum of oxidation states	+2	??	-14	0

- [a] potassium is a group one metal; its oxidation state is always +1
- [b] oxygen usually has an oxidation state of -2
- [c] the sum of oxidation states of each element is the product of the subscript and the oxidation state
- [d] find the -sum of the oxidation states of chromium (??) by setting the sum of all the oxidation states to zero

$$(+2) + ?? + (-14) = 0$$

$$?? = +12$$
- [f] find the oxidation state of chromium (?) by dividing the sum (+12) by the subscript (2)

$$+12 \div 2 = +6$$

Answer the questions below by circling the number of the correct response

- In this reaction, the oxidation number (oxidation state) of C changes from: $2\text{CO}_2 \rightarrow 2\text{CO} + \text{O}_2$
(1) 0 to +4 (2) +2 to +4 (3) +3 to 0 (4) +4 to +2
- In the reaction:
 $2\text{KMnO}_4 + 3\text{H}_2\text{SO}_4 + 5\text{H}_2\text{S} \rightarrow 5\text{S} + 2\text{MnSO}_4 + \text{K}_2\text{SO}_4 + 8\text{H}_2\text{O}$
the oxidation number of sulfur changes from
(1) +5 to -5 (2) -5 to +5 (3) 0 to -2 (4) -2 to 0
- What is the oxidation number of Cr in Na_2CrO_4 ?
(1) +1 (2) +2 (3) +3 (4) +6
- What is the oxidation state of the chromium in $\text{K}_2\text{Cr}_2\text{O}_7$?
(1) +5 (2) +6 (3) +3 (4) +12
- In the reaction $\text{Pb} + 2\text{Ag}^+ \rightarrow \text{Pb}^{2+} + 2\text{Ag}$, the reducing agent is
(1) Ag (2) Ag^+ (3) Pb (4) Pb^{2+}
- Which is not an oxidation-reduction reaction?
(1) $4\text{Na} + \text{O}_2 \rightarrow 2\text{Na}_2\text{O}$
(2) $\text{Fe} + 2\text{HCl} \rightarrow \text{FeCl}_2 + \text{H}_2$
(3) $\text{CaCl}_2(\text{aq}) + 2\text{AgNO}_3(\text{aq}) \rightarrow 2\text{AgCl}(\text{s}) + \text{Ca}(\text{NO}_3)_2(\text{aq})$
(4) $2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2$
- Given: $2\text{Al} + 3\text{Zn}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Zn}$. In this reaction, the oxidizing agent is (1) Al (2) Al^{3+} (3) Zn (4) Zn^{2+}
- Given: $2\text{Al} + 3\text{Zn}^{2+} \rightarrow 2\text{Al}^{3+} + 3\text{Zn}$. In this reaction, electrons are transferred from (1) Al to Al^{3+} (2) Zn^{2+} to Zn (3) Al to Zn^{2+} (4) Zn^{2+} to Al
- What is the oxidation number of nitrogen in N_2O_3 ? (1) +1 (2) +2 (3) +3 (4) +6
- In the reaction $3\text{CO} + \text{Fe}_2\text{O}_3 \rightarrow 3\text{CO}_2 + 2\text{Fe}$, the oxidation number of the iron changes from (1) +2 to 0 (2) +2 to +3 (3) +3 to +2 (4) +3 to 0
- What is the oxidation number of Br in BrO_3^{-2} ?
(1) +1 (2) +6 (3) +5 (4) +4
- Which is the reducing agent in the following reaction?
 $\text{Cl}_2(\text{aq}) + 2\text{KBr}(\text{aq}) \rightarrow 2\text{KCl}(\text{aq}) + \text{Br}_2(\text{aq})$
(1) Cl_2 (2) H_2O (3) K^+ (4) Br^-
- What is the oxidation number of carbon in $\text{C}_2\text{O}_4^{-2}$?
(1) +1 (2) +2 (3) +3 (4) +4
- Which is an oxidation-reduction reaction?
(1) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
(2) $\text{KOH} + \text{HBr} \rightarrow \text{KBr} + \text{H}_2\text{O}$
(3) $\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$
(4) $\text{Mg} + \text{Cl}_2 \rightarrow \text{MgCl}_2$
- MnSO_4 is a product in a reaction that contained KMnO_4 as a reactant. The oxidation number of the manganese changed from (1) -2 to +5 (2) +7 to +2 (3) +5 to -2 (4) -7 to +2
- Given the balanced equation:
 $2\text{HNO}_3 + 3\text{H}_2\text{S} \rightarrow 4\text{H}_2\text{O} + 2\text{NO} + 3\text{S}$
Which is reduced? (1) S (2) S-2 (3) N+2 (4) N+5
- During the reaction $\text{Ca} + \text{H}_2 \rightarrow \text{CaH}_2$, the oxidation number of the hydrogen changes from
(1) 0 to +1 (2) +1 to 0 (3) 0 to -1 (4) -1 to 0
- In the reaction $\text{Sn}^{+4} + \text{H}_2(\text{g}) \rightarrow \text{Sn}^{+2} + 2\text{H}^+$, the reducing agent is
(1) Sn^{+4} (2) H_2 (3) Sn^{+2} (4) H^+
- Given: $3\text{Ag} + 4\text{HNO}_3 \rightarrow \text{NO} + 3\text{AgNO}_3 + 2\text{H}_2\text{O}$. The reducing agent in this reaction is
(1) Ag (2) Ag^{+1} (3) H^{+1} (4) N^{+2}
- The reaction $\text{NaCl}(\text{s}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ is an example of
(1) an oxidation reaction, only
(2) a reduction reaction, only
(3) both an oxidation and a reduction reaction
(4) neither an oxidation nor a reduction reaction
- The oxidation number of manganese in KMnO_4 is
(1) +1 (2) +7 (3) +3 (4) +4
- In the reaction $\text{Sn}^{+2} + 2\text{Fe}^{+3} \rightarrow \text{Sn}^{+4} + 2\text{Fe}^{+2}$, the reducing agent is
(1) Fe^{+2} (2) Fe^{+3} (3) Sn^{+2} (4) Sn
- An oxidizing agent will always
(1) lose electrons (3) be reduced
(2) increase in oxidation number (4) increase in mass