Redox and Electrochemistry Test Review Sheet

Oxidation States (same as Oxidation Numbers)

Oxidation States (or Oxidation Numbers) are used to keep track of electrons in redox reactions and identify which species is oxidized or reduced. The term species is used to refer to the atom of an element.

Know the Rules for Assigning Oxidation Numbers:

1) Elements not combined with a different element have oxidation number of zero: Examples: Na Fe Cl_2 O_2

2) Ions have an oxidation number equal to the charge of the ion:

Examples of Ions alone: Na^{1+} Fe^{+2} Cl^{1-}

Examples of Ions in Ionic Compounds: NaCl Na in NaCl has a +1 oxidation state and Cl in NaCl has a -1 oxidation state

3) All group 1 elements have a +1 oxidation number as seen on the periodic table(except H). Similarly, group 2 must have +2. Al must be +3.

Ex: In KCl K must be +1 Cl is -1

In $MgCl_2$ Mg must be +2 each Cl is -1

*If only one ion/oxidation number is listed on the reference table - that is the oxidation number it MUST BE!

4) Oxygen has an oxidation number of **-2**. Unless it is in a peroxide (eg-H_2O_2), and in this case has an oxidation number of -1.

5) Hydrogen has an oxidation number of +1, unless it is bonded to a metal.

Ex: In HF, H=+1

In LiH, H = -1

6) The sum of oxidation numbers for a polyatomic ion must equal the ion's charge. (In a neutral compound, the sum of the oxidation #'s is 0)

CO_3^{2-} O is -2 but there are 3 so we have a total of (-6). C is +4

NH_4^+ H is +1 and there are 4 so we have (+4). N is -3

7	When neither oxygen nor hydrogen is present in a compound, the most electronegative atom is assigned the negative oxidation number (equal to the charge on the anion).	In SF_6, the oxidation number of each fluorine atom is –1. The oxidation number of the sulfur atom is +6 (+6 + (6 × –1) = 0).

Half –Reactions

LAW OF CONSERVATION OF CHARGE: THE CHARGE BEFORE A CHEMICAL OR PHYSICAL CHANGE EQUALS THE TOTAL CHARGE AFTER THE CHANGE. MAKE SURE THAT THE TOTAL CHARGE ON THE LEFT SIDE OF AN EQUATION EQUALS THE TOTAL CHARGE ON THE RIGHT SIDE OF THE EQUATION

Oxidation Half Reaction: Example: $Zn^0 \rightarrow Zn^{+2} + 2e^-$ (show electrons on right side)

Reduction Half Reaction: Example: $Cu^{+2} + 2e^- \rightarrow Cu^0$ (show electrons on left side)

Charge on both sides of the equation have to equal each other "Conservation of Charge"

DETERMINING WHETHER A "SPECIES" IS OXIDIZED OR REDUCED DEPENDS ON HOW THE OXIDATION STATE (OXIDATION NUMBER) CHANGES:

1. Oxidation State or Oxidation Number increases when electrons are lost (Oxidation)

2. Oxidation State or Oxidation Number decreases when electrons are gained (Reduction)

WHEN DIATOMIC ELEMENTS ARE IN A CHEMICAL EQUATION, KEEP THE DIATOMIC FORMULA IN THE HALF REACTION. EXAMPLE OF COMMON DIATOMIC ELEMENTS INCLUDE O_2 , Cl_2 , F_2 , N_2 , Br_2 , H_2 , I_2

Sample: Write the half reactions for the redox reaction: $Fe + Cl_2 \rightarrow FeCl_2$

Answer: Oxidation Half Reaction: $Fe^0 \rightarrow Fe^{+2} + 2e^-$

Reduction Half Reaction: $Cl_2^0 + 2e^- \rightarrow 2Cl^{1-}$

Balancing Redox Reactions

<u>PERFORM THE FOLLOWING STEPS</u>

1) Assign oxidation states

2) Identify element oxidized & reduced

3) Write the oxidation and reduction half reactions

4) Balance electrons gained & lost in half reactions and add the half reactions back together

5) Coefficients for the specific atoms in the balanced half reactions tells you how many of the atoms should be in the original unbalanced oxidation reduction equation

Ex:

$AlBr_3 + Fe \rightarrow FeBr_2 + Al$

1) Assign oxidation states

Al^{3+} $\cancel{3Br^-}$ Fe^0 Fe^{2+} $\cancel{2Br^-}$ Al^0

2) Identify element oxidized & reduced

Fe^0 was oxidized **Al^{3+} was reduced**

3) Write the oxidation and reduction half reactions

Oxidation half-reaction: $Fe \rightarrow Fe^{2+} + 2e^-$

Reduction half-reaction: $Al^{3+} + 3e^- \rightarrow Al$

4) Balance electrons gained & lost in half reactions and add the half reactions back together

Oxidation half-reaction: $(Fe \rightarrow Fe^{2+} + 2e^-) \times 3 = 3Fe \rightarrow 3Fe^{2+} + \cancel{6e^-}$

Reduction half-reaction: $(Al^{3+} + 3e^- \rightarrow Al) \times 2 = 2Al^{3+} + \cancel{6e^-} \rightarrow 2Al$

$3Fe + 2Al^{3+} \rightarrow 3Fe^{2+} + 2Al$

5) Coefficients for the specific atoms in the balanced half reactions tells you how many of the atoms should be in the original unbalanced oxidation reduction equation

$$2AlBr_3 + 3Fe \rightarrow 3FeBr_2 + 2Al$$

Spontaneous Reactions (USE TABLE J)

Metals	Nonmetals
Most Active ↓	Most Active ↓
Li	F_2
Rb	Cl_2
K	Br_2
Cs	I_2
Ba	
Sr	
Ca	
Na	
Mg	
Al	
Ti	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
H_2	
Cu	
Ag	
Au	
Least Active	Least Active

**Activity Series is based on the hydrogen standard. H_2 is *not* a metal.

Use Table J to predict or determine spontaneous reactions

For spontaneous reaction involving metals or hydrogen: **"species" oxidized must be more active in Table J than the element in Table J of the "species" reduced.** Usually the ion of the element in Table J is being reduced.

Example: $Mg + ZnCl_2 \rightarrow MgCl_2 + Zn$

Mg is oxidized and Zn^{+2} is reduced

Check Table J: Mg is more active than the element in Table J of species reduced (Zn^{+2}), therefore reaction is spontaneous

Example: $Cu + ZnCl_2 \rightarrow CuCl_2 + Zn$

Cu is oxidized and Zn^{+2} is reduced

Check Table J: Cu is not more active than the ion of the species reduced (Zn^{+2}), therefore the reaction is nonspontaneous

For- F_2, Cl_2, Br_2, I_2 - The most active species is the **most likely to be reduced (The opposite of the left column)**

Ex: $Cl_2 + 2NaBr \rightarrow NaCl + Br_2$ is nonspontaneous b/c Cl_2 is higher on Table J, so thus more likely to be reduced but is oxidized in this reaction.

Voltaic or Galvanic Cell (Battery)

Converts **chemical energy** into **electrical energy**

SPONTANEOUS redox reaction

Oxidation occurs at the the negative (-) Anode electrode (AN OX)

Reduction occurs at the positive (+) Cathode electrode (RED CAT)

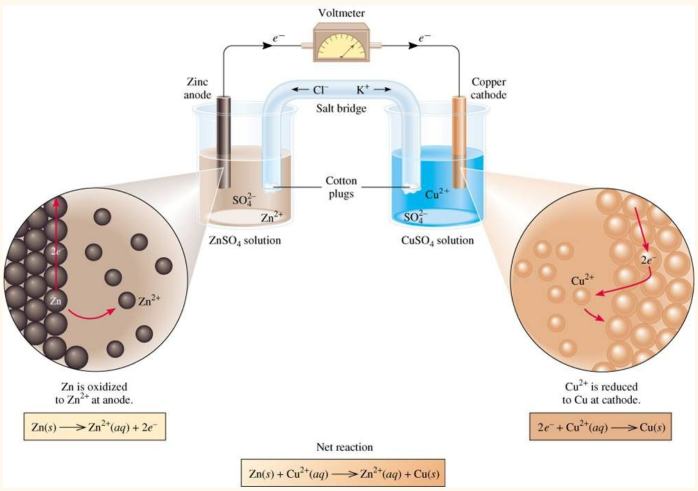
Purpose of the salt bridge is to allow ions to flow between half-cells to keep the solutions neutral.

Electrons flow from anode to cathode through the wire (NOT THROUGH THE SALT BRIDGE)

Batteries are examples of voltaic cells

As the REDOX reaction progresses, the anode loses mass and the cathode gains mass.

<u>Example of Voltaic Cell with a Salt Bridge</u>



Electrolytic Cell (Electrolysis)

Electrolysis is the process of using electricity to cause a chemical change.

Converts electrical energy into chemical energy.

NONSPONTANEOUS chemical reaction

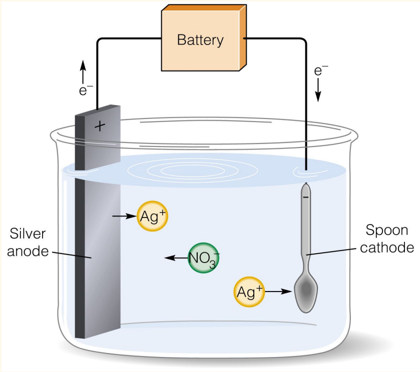
Oxidation occurs at the positive (+) Anode

Reduction occurs at negative (-) Cathode

Applications include: recharging a battery (reverse of spontaneous reaction) & electroplating

Electrolytic cell is not a battery, but uses electricity which could be provided by a battery

Example of an Electrolytic Cell: Electroplating



Voltaic vs. Electrolytic Cells Summary