Redox Reactions: Types of Reactions

Def: An oxidative-reduction (redox) reaction is a chemical reaction where a transfer of electrons occurs between 2 species.

Ex: $Cu^{+}_{(aq)} + Fe_{(s)} \longrightarrow Fe^{3+}_{(aq)} + Cu_{(s)}$

Electrons transfered from $Fe \to Cu$. How can we formalize this to reactions where this may not be obvious?

Def: Oxidation states/numbers describe the degree of oxidation of an atom in a chemical species. More practically, they are the hypothetical charge that an atom would have if all bonds to atoms of different elements were 100% ionic.

Remark: Oxidation states are typically integers ($\mathbb{Z}$), though rational oxidation states exist (as averages in a compound)

Remark: The sum of oxidation states in a compound should equal the overall charge of the compound. Neutral species have a total oxidation sum of 0. Ionic species sum to their charge.

Rules of Assigning Oxidation States:

1) Oxidation states of an individual atom is 0.

2) Group 1 elements have oxidation states of +1. Group 2 elements are +2.

3) The oxidation state of F is -1 in compounds

4) Hydrogen typically has a +1 oxidation number.

5) Oxygen typically has a +2 oxidation number.

Examples:
a) $Fe_{(s)}$, 0
b) $O_{2(g)}$, 0
c) $Fe_2O_{3(g)}$, Fe is +3 , O is -2

$2(+3) + 3(-2) = 0$!

Try it: i) $Cr_2O_7^{2-}$ ii) NO_3^-

Cr is +6 since
$2(+6) + 2(-7) = -2$

N is +5 since
$+5 + 3(-2) = 1$

So lets ~~[illegible]~~ examine the following reaction:

$$Fe_{(s)} + O_{2(g)} \longrightarrow Fe_2O_{3(g)}$$

The initial oxidation states are:
Fe : 0 } from before.
O : 0 }

Now, the product, Fe has +3 and Oxygen is -2.
So Fe lost electrons and Oxygen gained electrons. Motivates:

Def: During the course of a chemical reaction, if a compound or species loses electrons, we say it has been oxidized. If it gains electrons than it is said to be reduced.

So Fe was oxidized & O was reduced, during the course of the reaction.

Can use OIL RIG or LEO the lion says GER to remember.

Def: The compound or species that accepts electrons is called the oxidizing agent (causes other species to be ~~[illegible]~~ oxidized). Similarly, if it donates electrons, it is called the reducing agent (causes other species to be reduced).

So Fe oxidized but reducing agent.
And O reduced but oxidizing agent.

Okay, lets try it: identify the oxidation states of all elements and determine the oxidizing & reducing agent.

$$Fe_2O_{3(s)} + 2Al_{(s)} \longrightarrow Al_2O_{3(s)} + 2Fe_{(l)} \quad \text{(thermite rxn)}$$

Fe is +3 ⟶ 0
O is -2 ⟶ -2
Al is 0 ⟶ +3

So Fe gained e^- ⇒ reduced & thus oxidizing agent
Al lost e^- ⇒ oxidized & thus reducing agent

Now how can we balance redox reactions? From the begining we had

$$Cu^+_{(aq)} + Fe_{(s)} \longrightarrow Fe^{3+}_{(aq)} + Cu_{(s)}$$

This isn't balanced because the left has a +1 overall charge and the right has a +3 overall charge. We must have charges be equal (cannot create nor destroy charge!)

Well, we have

$$Cu^+_{(aq)} \longrightarrow Cu_{(s)} \text{ (gaining } e^-)$$

$$Fe_{(s)} \longrightarrow Fe^{3+}_{(aq)} \text{ (losing } e^-)$$

We can write these half-reactions:

$$Cu^+_{(aq)} + e^- \longrightarrow Cu_{(s)}$$
$$Fe_{(s)} \longrightarrow Fe^{3+}_{(aq)} + 3e^-$$

Want total charge conservation, so multiply Cu by 3 to get

$$3Cu^+_{(aq)} + 3e^- \longrightarrow 3Cu_{(s)}$$

Then adding the half-reactions:

$$3Cu^+_{(aq)} + 3e^- + Fe_{(s)} \longrightarrow 3Cu_{(s)} + Fe^{3+}_{(aq)} + 3e^-$$

The e^- cancel leaving:

$$\boxed{3Cu^+_{(aq)} + Fe_{(s)} \longrightarrow 3Cu_{(s)} + Fe^{3+}_{(aq)}}$$

What we just experienced was balancing half-reactions in a neutral solution. If we were in an acidic (H^+) or basic environment (OH^-), we would have to perform some additional steps. Don't worry about this right now!

Ex: In acidic conditions

$$Cr_2O_7^{2-}{}_{(aq)} + HNO_{2(aq)} \rightarrow Cr^{3+}{}_{(aq)} + NO_3^-{}_{(aq)}$$

Start with half-rxns:

$Cr_2O_7^{2-}{}_{(aq)} \longrightarrow 2Cr^{3+}{}_{(aq)}$ $\qquad$ $HNO_{2(aq)} \rightarrow NO_3^-{}_{(aq)}$

1) Add H_2O to balance O's.

$Cr_2O_7^{2-}{}_{(aq)} \rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$ $\qquad$ $HNO_{2(aq)} + H_2O_{(l)} \rightarrow NO_3^-{}_{(aq)}$

2) Balance H's via H^+ (acidic environment).

$$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} \rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)}$$

$$HNO_{2(aq)} + H_2O_{(l)} \rightarrow NO_3^-{}_{(aq)} + 3H^+_{(aq)}$$

3) Balance charges using e^-:

$$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} \rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)} + 6e^-$$

$$HNO_{2(aq)} + H_2O_{(l)} + 2e^- \rightarrow NO_3^-{}_{(aq)} + 3H^+_{(aq)}$$

4) Multiply so e^- cancel and add:

$$Cr_2O_7^{2-}{}_{(aq)} + 14H^+_{(aq)} + 3HNO_{2(aq)} + 3H_2O_{(l)}$$
$$\rightarrow 2Cr^{3+}{}_{(aq)} + 7H_2O_{(l)} + 3NO_3^-{}_{(aq)} + 9H^+_{(aq)}$$

5) Cancel/simplify

$$\boxed{3HNO_{2(aq)} + Cr_2O_7^{2-}{}_{(aq)} + 5H^+_{(aq)} \rightarrow 3NO_3^-{}_{(aq)} + 2Cr^{3+}{}_{(aq)} + 4H_2O_{(l)}}$$

So what we have talked about so far are redox reactions, but we can classify other types of reactions

i) Precipitation

$$BaCl_{2(aq)} + K_2CO_{3(aq)} \rightarrow BaCO_{3(s)} + 2KCl_{(aq)}$$

The $BaCO_3$ is insoluble in water & falls out of solution. Easy to identify precipitation type reactions since they will form a solid from liquid or aqueous reagents.

ii) Acid-Base

$$HNO_{3(aq)} + KOH_{(aq)} \rightarrow H_2O_{(l)} + KNO_{3(aq)}$$

Usually identified by formation of water + salt. What is given above is a neutralization type reaction. Also have hydration of acid/base anhydrides to form acid/bases

$$CaO_{(aq)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)}$$

and acid/base displacement:

$$\underbrace{H_2SO_{4(aq)}}_{\text{acid}} + 2NaCl_{(aq)} \rightarrow Na_2SO_{4(aq)} + \underbrace{2HCl_{(aq)}}_{\text{acid}}$$

iii) Gas Evolution (Not Part of lab)

$$NH_4Cl_{(aq)} + CsOH_{(aq)} \rightarrow \boxed{NH_{3(g)}} + H_2O_{(l)} + CsCl_{(aq)}$$

Formation of gas (hence the name).