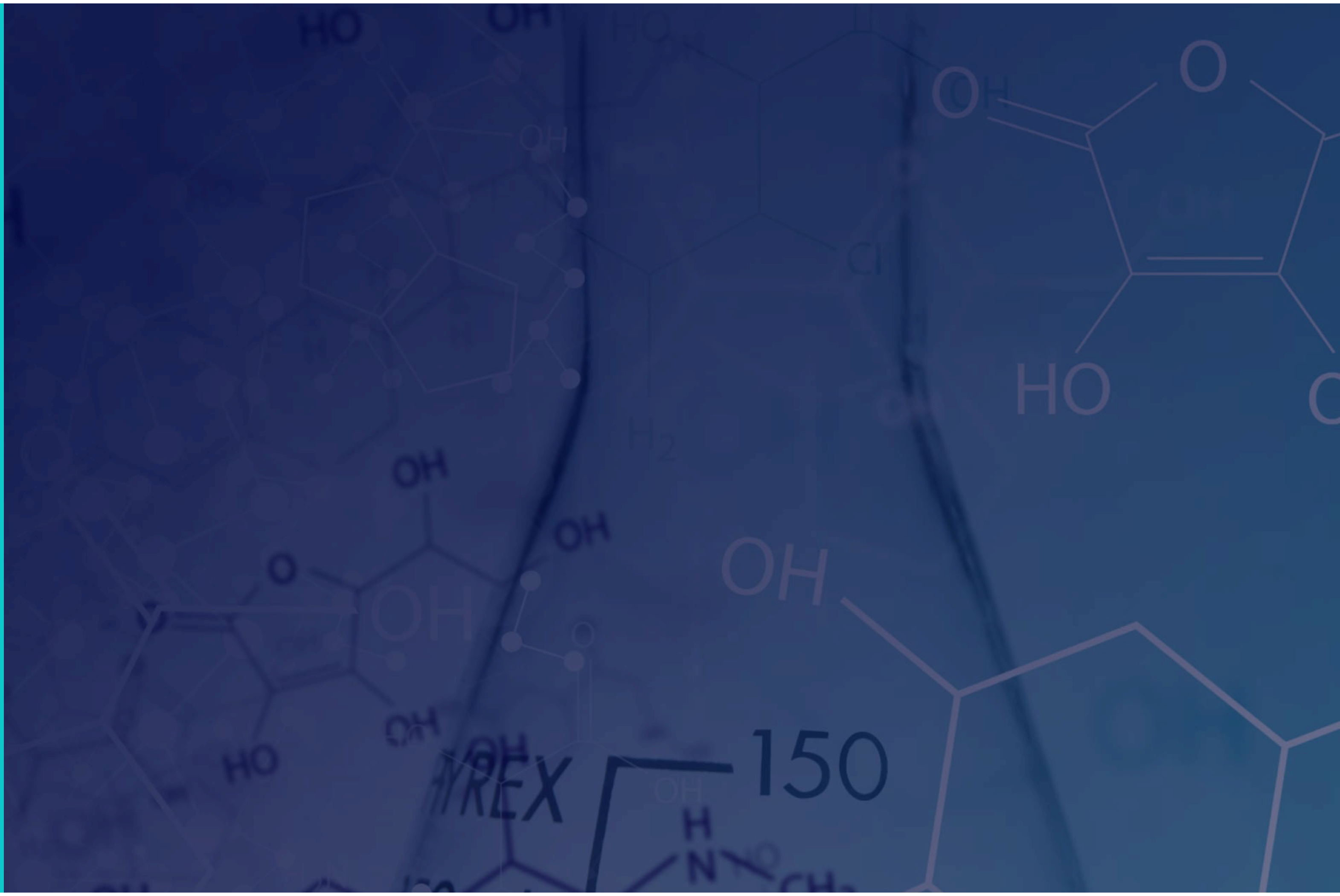

UWO

CHEM 1302

Winter 2025, Chapter 11 Notes

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11. Redox Reactions

11.1 Oxidation States

11.1.1

Oxidation States

Oxidation state: A way to describe the degree of oxidation of a chemical

WIZE TIP

Tips to Determine Oxidation State:

1. Free elements: oxidation state = 0
Examples: $Cl_2(g)$ has an oxidation state of 0, Na(s) has an oxidation state of 0

2. Monoatomic ions: oxidation state = ionic charge
Examples: $Cl^-(aq)$ has an oxidation state of -1, $Na^+(aq)$ has an oxidation state of +1.

3. Hydrogen: oxidation state = +1
   a. **Exception: hydrogen in hydrides** (e.g., NaH) – **oxidation state = -1**

4. Oxygen: oxidation state = -2
   a. **Exception: oxygen in peroxides** (e.g., H_2O_2) – **oxidation state = -1**

5. For **neutral molecules, oxidation numbers add up to zero.**
For **polyatomic ions, oxidation numbers add up to the ion charge.**
Examples: $Cl_2(g)$ has an oxidation state of 0, NaCl oxidation numbers add up to 0, SO_4^{2-} oxidation numbers add up to 2-.

11.1.2

Oxidation States - Additional Information

Other Low Yield Tips to Help Determine Oxidation States

- **Alkali metals: oxidation state= +1**
- **Alkaline earth metals: oxidation state= +2**
- **Nitrogen Group: oxidation state= -3**
- **Oxygen family: oxidation state= -2**
- **Halogens: oxidation state= -1**

Group → / ↓Period	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba		72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra		104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Nh	114 Fl	115 Ms	116 Lv	117 Ts	118 Og
Lanthanides			57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	
Actinides			89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Note on Transition Metals

Transition metals can have different charges (ex. Pb^{2+} and Pb^{4+})
As a result, transition metals can also have different oxidation states!

Example:

What is the oxidation state of Cu in CuCl? ______

What is the oxidation state of copper in $Cu(NO_3)_2$? ______

When naming compounds involving transition metals, the **oxidation state of the transition metal is written in Roman numerals and put in brackets:**

- copper(I) chloride: CuCl
- copper(II) nitrate: $Cu(NO_3)_2$

Example: Oxidation State Practice

What is the oxidation state of:

S in S_2O_3: ______

O in S_2O_3: ______

What about Pb in Pb_2O_3?

This one is different because this compound is a mixed oxidation state compound, so although Pb should have 2+ or 4+, here, we get an oxidation state of ______ and that is ok because both Pb^{2+} and Pb^{4+} are present in the compound.

11.1.4

What is the oxidation state of Br in BrO_3^- ?

- +4
- +5
- +6
- -4
- -5
- -6

11.1.5

What is the oxidation state for the following: (enter answer as -1 or +2 for example)

a) P in P_2O_5:

b) Cr in $Na_2Cr_2O_7$:

c) Al in $Al_2(SO_4)_3$:

d) Na in Na_2O_2:

e) O in Na_2O_2:

f) N in NH_3:

11.2 Reduction-Oxidation (Redox) Reactions

11.2.1

Redox Reactions

OIL RIG! Oxidation Is Loss of electrons, Reduction Is Gain of electrons
OR
LEO GER! Loss of Electrons is Oxidation, Gain of Electrons is Reduction

OXIDATION (lose e-)	REDUCTION (gain e–)
Reducing agent reduces others and is itself oxidized.	**Oxidizing agent** oxidizes others and is itself reduced.
Oxidation number increases.	Oxidation number is reduced.

Keep in mind that in organic chemistry the definitions of oxidation and reduction are different:

- For **reduction**, we would look for **more bonds to H** and **less bonds to O**
- For **oxidation**, we would look for **less bonds to H** and **more bonds to O!**

How can you recognize if a redox reaction is occuring?

1. Can **check oxidation #s of each element** to see if they **change from reactants to products** (most important for now!)

2. Familiarize yourself with **common oxidizing and reducing agents** (you'll learn more about this in orgo!)

- **Oxidizing agents:** O_2, halogens, HNO_3, Cr_2O_72-, MnO_4-
- **Reducing agents:** H_2, metals, C

11.2.2

Practice: Oxidation vs Reduction

Consider the following balanced redox equation:

$$16H^{+}_{(aq)} + 2MnO^{-}_{4(aq)} + 10Cl^{-}_{(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 5Cl_{2(g)} + 8H_2O_{(l)}$$

Which species is being oxidized?

- H^+
- MnO_4^-
- Cl^-
- Mn^{2+}
- Cl_2

11.2.3

Practice: Oxidizing vs Reducing Agent

Consider the following balanced redox equation:

$$16H^+_{(aq)} + 2MnO^-_{4(aq)} + 10Cl^-_{(aq)} \rightarrow 2Mn^{2+}_{(aq)} + 5Cl_{2(g)} + 8H_2O_{(l)}$$

Which species is the reducing agent?

- H^+
- MnO_4^-
- Cl^-
- Mn^{2+}
- Cl_2

11.2.4

Practice: Oxidizing vs Reducing Agents

What is the reducing agent in the following chemical reaction?

$NO + Au^{+} \rightarrow NO_3^{-} + Au$

- Au^{+}
- NO
- NO_3^{-}
- Au

11.3 Balancing Redox Reactions [half-reaction method]

11.3.1

Balancing Redox Reactions: Half-Reaction Method

The "half-reaction method" of balancing redox reactions is the most commonly used method :)

Balance the following redox reaction in basic conditions and assign oxidation numbers to all species.

$$\mathrm{Ag\,(s) + \ Zn^{2+}\,(aq)\ \rightarrow\ Ag_2O\,(aq)\ + Zn\,(s)}$$

WIZE TIP

Balancing redox equations:

1. **Check the oxidation numbers** of each species by inspection (**change in oxidation # means this is a redox reaction!**)

2. **Write** the reduction and oxidation **half-reactions.**

3. **Balance each** half-reaction: start with **atoms that are neither O nor H, balance oxygens with H_2O, hydrogens with H+, and remaining charges with e–.**

4. **Multiply** the reduction and oxidation half-reactions as needed so that each reaction exchanges the **same number of e–.**

5. **Add** the oxidation and reduction half reactions together and simplify.

6. **Additional step in basic solution only:** there can't be any H+ ions, so **neutralize the H+ by adding the same # of OH- to each side.**

1. Check the oxidation numbers of each species by inspection (change in oxidation # means this is a redox reaction!)

Here we are told it's a redox reaction already, but we can still check oxidation #s to see what is oxidized and what is reduced:

$$\mathrm{Ag\,(s) + \ Zn^{2+}\,(aq) \ \rightarrow \ Ag_2O\,(aq) \ + Zn\,(s)}$$

- Ag goes from having an oxidation # of _____ to ______. Therefore Ag is ___________.
- Zn goes from having an oxidation # of ______ to _____. Therefore Zn is ___________.

2. Write the reduction and oxidation half-reactions.

Oxidation Half Reaction:

Reduction Half Reaction:

3. Balance each half-reaction: start with atoms that are neither O nor H, balance oxygens with H_2O, hydrogens with H+, and remaining charges with e−.

a) Balance everything that is not O and H:

ox:

red:

ox: $2Ag(s) \rightarrow Ag_2O(aq)$

red: $Zn^{2+}(aq) \rightarrow Zn(s)$

b) Balance O with H_2O:

ox:

red:

c) Balance H with H+

ox:

red:

d) Balance charges with electrons!

ox:

red:

ox:

$2Ag(s) + H_2O \rightarrow Ag_2O(aq) + 2H^+(aq) + 2e^-$

red:

$Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$

4. Multiply the reduction and oxidation half-reactions as needed so that each reaction exchanges the same number of e−.

5. Add the oxidation and reduction half reactions together and simplify.

$$2\mathrm{Ag}\,(\mathrm{s}) + \ \mathrm{H_2O} + \mathrm{Zn}^{2+}(aq) \rightarrow \mathrm{Zn(s)} + \mathrm{Ag_2O}\,(\mathrm{aq}) \ + 2\mathrm{H}^{+}(\mathrm{aq})$$

6. Additional step in basic solution only: there can't be any H+ ions, so neutralize the H+ by adding the same # of OH- to each side.

11.3.2

Practice: Balancing Redox Reactions

Balance the following redox reaction in an acidic solution.

$$HCOOH + MnO_4^- \text{-->} CO_2 + Mn^{2+}$$

11.4 Reduction Potentials

11.4.1

Reduction Potentials

- Each half reaction has a potential.
- For example, the first half reaction below has a reduction potential of -3.05V!
  - We got these potentials from measuring the voltage when the other half reaction was the **reference electrode (H 0V)**
- In the table below, are the reactions showing reductions or oxidations? ____________
  - Therefore, these are ____________ potentials!
- **If we flip a reaction, the E^o potential changes signs**, let's do this for the first half reaction in the table so it is clear:

- Is this reaction now describing reduction or oxidation? ____________!
  - So do you think the potential written would now be a reduction or oxidation potential? ____________ potential
- **In general:**
  - **If we see a more positive reduction potential:**
    - It means something is more likely to be oxidized/reduced: ____________
    - and will act as a(n) reducing/oxidizing agent: ____________ agent
  - **If we see a more negative reduction potential:**
    - It means something is more likely to be oxidized/reduced: ____________
    - and will act as a(n) reducing/oxidizing agent: ____________ agent

Half Reaction	Potential E°
$Li^{+}(aq) + e^{-} \dashrightarrow Li(s)$	-3.05
$Mg^{2+}(aq) + 2e^{-} \dashrightarrow Mg(s)$	-2.36
$Al^{3+}(aq) + 3e^{-} \dashrightarrow Al(s)$	-1.67
$H_2O(l) + 2e^{-} \dashrightarrow H_2(g) + 2OH^{-}(aq)$	-0.83
$Zn^{2+}(aq) + 2e^{-} \dashrightarrow Zn(s)$	-0.76
$Fe^{2+}(aq) + 2e^{-} \dashrightarrow Fe(s)$	-0.44
$Ni^{2+}(aq) + 2e^{-} \dashrightarrow Ni(s)$	-0.23
$Pb^{2+}(aq) + 2e^{-} \dashrightarrow Pb(s)$	-0.13
$Fe^{3+}(aq) + 3e^{-} \dashrightarrow Fe(s)$	-0.036
$2H^{+}(aq) + 2e^{-} \dashrightarrow H_2(g)$	0
$Cu^{2+}(aq) + 2e^{-} \dashrightarrow Cu(s)$	0.34
$Cu^{+}(aq) + e^{-} \dashrightarrow Cu(s)$	0.52
$Hg^{2+}(aq) + 2e^{-} \dashrightarrow Hg(l)$	0.8
$Ag^{2+}(aq) + 2e^{-} \dashrightarrow Ag(s)$	0.8
$Pt^{2+}(aq) + 2e^{-} \dashrightarrow Pt(s)$	1.2
$O_2(g) + 4H^{+}(aq) + 4e^{-} \dashrightarrow H_2O(l)$	1.23
$Au^{3+}(aq) + 3e^{-} \dashrightarrow Au(s)$	1.5
$F_2(g) + 2e^{-} \dashrightarrow 2F^{-}(aq)$	2.87

- Based on this table, answer the following questions:

1) What is the strongest oxidizing agent? ________

2) What is the weakest reducing agent? ________

3) What is the strongest reducing agent? ________

WATCH OUT!

DO NOT assume you are given REDUCTION potentials in the table!
Here we know they are reduction potentials because we see electrons being added on the left sides of the equations.
Your prof may try to trick you and show you oxidation potentials instead, where electrons would be seen on the right sides of the equations.

Another thing we have to know is that we may need to multiply half reactions by a coefficient (later on in problems). When we do this NEVER multiply the potential by a coefficient.
For example in the table we have:
$Li^{+}(aq) + e^- \rightarrow Li(s)$ $E^{o} = -3.05V$

If I needed to multiply this half reaction by 2, I would get:
$2Li^{+}(aq) + 2e^- \rightarrow 2Li(s)$ $Eo = -3.05V$

How Potentials Relate to Spontaneity

- The overall cell potential will determine the spontaneity of redox reactions:

$$\Delta G^o = -nFE^o_{redox}$$

E^o_{redox}=the potential associated with the redox reaction

n=# of electrons

F=Faraday's constant (96 500C/mol e-)

deltaG^o=Gibbs free energy

- According to this equation, **when E^o_{redox} is positive**....deltaG^o is positive or negative:

  ____
  - This means the reaction would be spontaneous/non-spontaneous: ________________
  - When this is the case we have a **galvanic cell** (more on this soon!)
  - K _____ 1
- **When E^o_{redox} is negative**...deltaG^o is positive or negative: ___________
  - This means the reaction would be spontaneous/non-spontaneous: ____________________
  - When this is the case we have an **electrolytic cell** (more on this soon!)
  - K _____ 1

Apply: If we had the following overall reaction, would it be spontaneous or not?
$Zn(s) + Cu^{2+}(aq) \dashrightarrow Zn^{2+}(aq) + Cu(s)$

1) Let's write out the oxidation and reduction half reactions:

a) Oxidation half reaction:

b) Reduction half reaction:

2) Given the above reactions that are occuring, write out an oxidation potential and a reduction potential using the table values below:

From the table:

Zn^{2+} (aq) + 2e- -> Zn(s) Eo=-0.76V

Cu^{2+}(aq) + 2e- -> Cu(s) Eo=+0.34V

For us, _________ is being oxidized, so E^{o}_{ox}= _________ (changes sign)

And ____________ is being reduced, so E^{o}_{red}= __________ (stays the same)

3) Solve for E^{o}_{redox}:

$$E^{o}_{redox} = E^{o}_{ox} + E^{o}_{red}$$

E^{o}_{redox}= 0.76 + 0.34

E^{o}_{redox}=1.10V

4) Figure out if the reaction is spontaneous using this equation: $\Delta G^{o} = -nFE^{o}_{redox}$

- Our E^{o}_{redox} value is positive/negative: ____________
- This means our deltaG^{o} value is positive/negative: ____________
- Therefore, the reaction is spontaneous/non-spontaneous: ________________

11.4.3

Will the following reaction proceed spontaneously or will current need to be added externally?

$$Cu_{(s)} + Ca^{2+}_{(aq)} \rightarrow Ca_{(s)} + Cu^{2+}_{(aq)}$$

Standard reduction potentials

$$\begin{array}{ll} Ca^{2+}_{(aq)} + 2e^{-} \rightarrow Ca_{(s)} & E^{o} = -2.868V \\ Cu^{2+}_{(aq)} + 2e^{-} \rightarrow Cu_{(s)} & E^{o} = -0.340V \end{array}$$

-
-
-
-

11.4.4

Based on the following reduction potentials, which of these species is the strongest reducing agent?

Reaction	$E^\circ(V)$
$2HC\ell O(aq) + 2H^+(aq) + 2e^- \rightarrow C\ell_2(g) + 2H_2O(\ell)$	$+1.63$
$C\ell_2(g) + 2e^- \rightarrow 2C\ell^-(aq)$	$+1.36$
$Br_2(\ell) + 2e^- \rightarrow 2Br^-(aq)$	$+1.08$
$I_2(s) + 2e^- \rightarrow 2I^-(aq)$	$+0.535$
$Ag + (aq) + e^- \rightarrow Ag(s)$	$+0.799$
$Sn^{4+}(aq) + 2e^- \rightarrow Sn^{2+}(aq)$	$+0.15$

- I_2 (s)
- I^- (aq)
- Br_2 (ℓ)
- Br^- (aq)
- Ag(s)
- Sn^{2+} (aq)
- Sn^{4+} (aq)