Chemistry 12 – Unit 5

Oxidation – Reduction

Introduction

-Demonstration of oxidation – reduction reactions

Definitions: (**species** means atom, ion or molecule)

Oxidation – a species undergoing oxidation **loses electrons**
(charge becomes more **positive**)

Reduction – a species undergoing reduction **gains electrons**
(charge becomes more **negative**)

Oxidizing agent – The **species being reduced**
(gains electrons, causes the other one to be oxidized)

Reducing agent – The **species being oxidized**
(loses electrons, causes the other one to be reduced)

$2\ e^-$

E.g.) $Cu^{2+}_{(aq)} + Zn_{(s)} \rightarrow Cu_{(s)} + Zn^{2+}_{(aq)}$

Oxidizing agent

LEO says GER	
Losing **E**lectrons is **O**xidization	**G**aining **E**lectrons is **R**eduction

OAR
The **O**xidizing **A**gent is **R**educed

To carry it too far…

When LEO the Lion says GER you grab your OAR and Row Away Outa' there!
(Reducing Agent is Oxidized)

Redox – Short for **Oxidation – Reduction**

Redox identification

Charge on neutral atom or molecule = 0

Oxidation – Charge gets more **+** (loses electrons)

Reduction – Charge gets more **–** (gains electrons)

Reduction (charge decreases)

E.g.) $Pb^{2+}_{(aq)} + Mg^{0}_{(s)} \rightarrow Pb^{0}_{(s)} + Mg^{2+}_{(aq)}$

Oxidation (Charge increases)

Question

In the reaction:

$2Fe^{2+} + Cl_2 \rightarrow 2Fe^{3+} + 2Cl^-$

Identify:

a) The Oxidizing Agent: ____________________
b) The species being oxidized:________________
c) The reducing agent:_____________________
d) The species being reduced:________________
e) The species gaining electrons:______________
f) The species losing electrons:_______________
g) The product of oxidation__________________
h) The product of reduction__________________

Do Ex. 1 (a-e) pp. 192 SW

Half-Reactions

-Redox reactions can be broken up into oxidation & reduction half reactions.

e.g.) Redox rx: $Pb^{2+}_{(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)}$

The Pb^{2+} (loses/gains) __________ 2 electrons.

Reduction Half-rx: $Pb^{2+}_{(aq)} + 2e^- \rightarrow Pb_{(s)}$

Electrons on the LEFT side (or GER)
Means REDUCTION

Write the **oxidation** half reaction for the following redox rx.

$Pb^{2+}_{(aq)} + Zn_{(s)} \rightarrow Pb_{(s)} + Zn^{2+}_{(aq)}$

Ox half rx: ____________________________
(In oxidation reactions, e^-'s are ____ and are found on the ____ side.) (LEO)

Note: Half-rx's always have e^-'s, redox (oxidation-reduction) reactions never show e^-'s!

Given the redox reaction:

$F_{2(g)} + Sn^{2+}_{(aq)} \rightarrow 2F^-_{(aq)} + Sn^{4+}_{(aq)}$

Write the **oxidation** half-rx:___________________________

Write the **reduction** half-rx:___________________________

Do ex. 2 a-c on p. 192 SW

Oxidation numbers

-Real or apparent charge on an atom in a molecule or ion

In SW. p. 193 -the charge that an atom would possess if the species containing the atom was made up of ions (even if it's not!)

Rules to find oxidation number of an atom

1) **In elemental form**:

(Single atoms of monatomic elements) or (diatomic molecules of diatomic elements)

Oxidation number of atoms = 0

Eg) $\underbrace{\text{Mn, Cr, } N_2\text{, } F_2\text{, Sn, } O_2\text{, etc.}}$

The oxidation # of each atom = 0

2) In **monatomic ions**: oxidation # = charge

Eg) In Cr^{3+} -oxidation # of Cr = +3

S^{2-} -oxidation # of S = -2

3) In **<u>ionic compounds</u>**

a) the oxidation # of **Alkali Metals** is always **+1**

eg) $\underline{Na}Cl$ $\underline{K}_2CrO_4$

Ox # of Na & K = +1

b) the oxidation # of **Halogens** when at the end (right side) of the formula is always **–1**

eg) $Ca\mathbf{Cl}_2$ $Al\mathbf{Br}_3$ $K\mathbf{F}$

Ox # of Cl, Br and F = -1

Note: Halogens are **not** always –1! (Only when it is written **last** in formula.)

4) <u>In **molecules** or **polyatomic ions**:</u>

a) Ox. # of **oxygen** is almost always **–2**

e.g.) $\underbrace{K\underline{O}H \quad Cr\underline{O}_4^{2-} \quad Li_3P\underline{O}_4}_{\text{Ox # of O is –2}}$

b) An exception is **<u>Peroxides</u>** in which ox. # of O = **-1**

<u>Hydrogen Peroxide:</u> $H_2\mathbf{O}_2$

<u>Alkali Peroxides:</u> $Na_2\mathbf{O}_2$

Ox # of O's = -1

(Remember, "O" in O_2 has an Ox. # of ______)

5) <u>In **molecules** or **ions**:</u>

a) **Hydrogen** is almost always **+1**

e.g.) $\underline{H}NO_3$ $\underline{H}_2SO_4$ $\underline{H}PO_4^{2-}$ } Every "H" has an ox # of +1

b) An exception is **metallic hydrides**, which have an ox # of **-1**

e.g.) $Na\underline{H}$ $Ca\underline{H}_2$ (In each one of these Ox. # of H = -1)

(What is the ox # of "H" in NH_3? _______)

(And remember ox # of "H" in H_2 = ______)

<u>Finding oxidation numbers of each atom in a molecule or PAI</u>

In a **neutral molecule** the **total charge = 0**

e.g.) NH_3 ← Total charge = 0 (no charge)

In a **polyatomic** ion – the **total ionic charge** is written on the **top right**

e.g.) CrO_4^{2-} ← Total ionic charge (TIC) = -2

<u>Oxidation numbers</u> <u>of all atoms</u> **<u>add up to total ionic charge (TIC)</u>**

e.g.) Find the oxidation # of Cr in CrO_4^{2-}

(Let x = ox # of one Cr atom)

$\underline{Cr}O_4^{2-}$

X + 4 [# of "O"atoms] (-2 [charge of oxygen]) = -2 [total ionic charge]

X – 8 = -2

X = -2 + 8

X = **+6** So ox # of Cr here = +6

e.g.) Find ox # of Cl in $HClO_4$

$HClO_4$

+1 + x + 4 (-2) = 0

1 + x – 8 = 0

x – 7 = 0

x = **+7**

e.g.) Find Ox # of Cr in $Cr_2O_7^{2-}$

$Cr_2O_7^{2-}$

2x + 7(-2) = -2

2x – 14 = -2

2x = +12

x = **+6**

e.g.) Find ox # of P in Li_3PO_4

$Li_3\ P\ O_4$

$3(+1) + x + 4\ (-2) = 0$

$3 + x - 8 = 0$

$x - 5 = 0$

$x = \mathbf{+5}$

Find Ox # of the underlined element in each of the following:

a) $NaH_2\underline{P}O_4$ ______ b) $Na_2\underline{O_2}$ _______ c) $K\underline{H}$ _______

Find the ox # of **Fe** in Fe_3O_4

Find the ox # of **As** in As_3O_5

Read p. 193-194 of SW. Do Exercise 3 on p. 194 of SW.

Changes in oxidation numbers

When an atom's **oxidation # is <u>increased</u>**, it is **<u>oxidized</u>**.

e.g.) Half-rx: $Fe^{2+} \rightarrow Fe^{3+} + e^-$

<u>More complex:</u>

-When Mn^{3+} changes to MnO_4^-, is Mn oxidized or reduced?

$Mn^{3+} \rightarrow MnO_4^-$

- What is the ox # of Mn before & after the reaction? Before ___ After ___

- The ox # of Mn is (de/in)____creased.

- In this process, Mn is (*oxidized/reduced*)__________________

Reduction – When an atom's **oxidation # is decreased**, it is **reduced**.

e.g.) $Cu(NO_3)_2 \rightarrow Cu_{(s)}$ Ox # **decreases** (**reduction**)

Ox # of Cu = +2

Ox # of Cu = 0

Redox ID using oxidation #'s

Given a more complex equation – identify atoms which **do not change ox #'s** (often "O" or "H" but not always!)

e.g.) $3SO_2 + 3H_2O + ClO_3^- \rightarrow 3SO_4^{2-} + 6H^+ + Cl^-$

There are no O_2 molecules or peroxides, so "O" in all these has an ox # = -2

$3SO_2 + 3H_2O + ClO_3^- \rightarrow 3SO_4^{2-} + 6H^+ + Cl^-$

H is (+1) in both of these so it doesn't change

Again:

$3SO_2 + 3H_2O + ClO_3^- \rightarrow 3SO_4^{2-} + 6H^+ + Cl^-$

The only atoms left are "S" and "Cl". Find the Ox #'s of S and Cl^- in species that contain them. (Ox # of 1 atom in each case)

$3SO_2 \rightarrow 3SO_4^{2-}$

Coefficients are just for balancing.

$SO_2 \rightarrow SO_4^{2-}$

Ox # of S is +4 Ox # of S is +6

Ox # of S **increases** so S is being **oxidized**

Note:

- R.A.O., the **reducing agent** is **oxidized**
- The species **SO_2** is acting as the **reducing agent**.
- The element **S** is being oxidized so **S is losing electrons**.

Look at the species with Cl:

ClO_3^- → Cl^-

Ox. # of Cl = **+5**	Ox. # of Cl = **-1**

Decrease in ox # so Cl is being **reduced**

Therefore, the **species** acting as the **oxidizing agent** is __________.

(They may also ask for the **atom** acting as the **oxidizing agent**
– this would be **Cl** in ClO_3^-)

Eg. –given the reaction:

$2CrO_4^{2-} + 3HCHO + 2H_2O \rightarrow 2Cr(OH)_3 + 3HCOO^- + OH^-$

Find: a) The species being oxidized
c) The reducing agent
d) The species being reduced
e) The oxidizing agent
f) The species losing electrons
g) The species gaining electrons

Notes:
- For hydrocarbons it's best to rewrite them as simple molecular formulas.
- All O's are in molecules or ions, no O_2 & no peroxides so O remains unchanged as -2
- All H's are in molecules or ions, no H_2 or metallic hydrides so H remains unchanged as +1
- The atoms to check for changes are C and Cr.

$0 \xrightarrow{\text{Oxidation}} +2$

$2\underline{Cr}O_4^{2-} + 3\underline{C}H_2O + 2H_2O \rightarrow 2\underline{Cr}(OH)_3 + 3H\underline{C}O_2^- + OH^-$

$+6 \xrightarrow[\text{Reduction}]{} +3$

So… a) the species being **oxidized** is (CH_2O) HCHO (inc. in ox #)
b) the **reducing agent** is (CH_2O) HCHO (RAO)
c) The species being **reduced** is CrO_4^{2-}(decrease in ox #)
d) The **oxidizing agent** is CrO_4^{2-} (OAR)
e) The species **losing e^-'s** is (CH_2O) HCHO (LEO)
f) The species **gaining e^-'s** is CrO_4^{2-} (GER)

Given the redox reaction:

$2MnO_4^- + 3C_2O_4^{2-} + 4H_2O \rightarrow 2MnO_2 + 6CO_2 + 8OH^-$

Find:

a) The species being reduced: ______________.
b) The species undergoing oxidation: ______________.
c) The oxidizing agent: ________________.
d) The reducing agent: ________________.
e) The species gaining electrons: ______________.
f) The species losing electrons: _______________.

Given the balanced redox reaction:

$3S + 4HNO_3 \rightarrow 3SO_2 + 4NO + 2H_2O$

Find:

a) The oxidizing agent: ______________.
b) The reducing agent: ______________.
c) The species being reduced: ____________.
d) The species being oxidized: ____________.
e) The species losing electrons: _____________.
f) The species gaining electrons: _____________.
g) The product of oxidation: _____________.
h) The product of reduction: _____________.

Given the following:

$6Br_2 + 12KOH \rightarrow 10KBr + 2KBrO_3 + 6H_2O$

Find:

a) The oxidizing agent: _______________.
b) The reducing agent: _______________.
c) The species undergoing oxidation: ______________.
d) The species being reduced: _________________.
e) The product of oxidation: ________________.
f) The product of reduction: ________________.

Using oxidation numbers to identify half-reactions

They don't have to be balanced

e.g.) If $NO_2^- \rightarrow NO_3^-$ is an example of (oxidation or reduction?) ________________.

("O" does not change it's ox # (no O_2 or peroxides)) so find ox # of **N** on both sides.

$NO_2^- \rightarrow NO_3^-$

Ox # = +3 Ox # = +5

$\triangle$ O.N. = +2

Since ox # increases, this is an *oxidation* e.g.) $H_2\mathbf{O}_2$ → $H_2\mathbf{O}$

Ox # = -1 Ox # = -2

△ O.N. = -1

This time "O" is the element changing ox #

Ox # if "O" **decreases** so this is a **reduction**

<u>Find the ∆ O.N. of the element in which it changes and identify each as an oxidation or reduction</u>

a) $C_2H_5OH \rightarrow CH_3COOH$ ________________________________
b) $Fe_2O_3 \rightarrow Fe_3O_4$ ______________________________________
c) $H_3PO_4 \rightarrow P_4$ ______________________________________
   (P_4 is the elemental form of phosphorus)
d) $CH_3COOH \rightarrow CH_3COH$ ________________________________

NOTE: When asked if a given reaction is a redox or not:

Look for a change from **an element → compound** or **compound→ an element**

These will **<u>always</u>** be redox, because in **elemental form ox. # = 0** and in compounds usually **ox. # is not = 0**

$$\overset{0 \longrightarrow +2}{}$$
Eg.) Is the reaction: $Zn + Cl_2 \rightarrow ZnCl_2$ a redox reaction?
$$\underset{0 \longrightarrow -1}{}$$

Answer: It must be because ∆ON of Zn (0 → +2 = +2) and ∆ON of Cl (0 → -1 = -1)

Do Exercises 4, 5 and 6 on p. 194-195 of SW.

<u>Half-reactions and the reduction table</u>

<u>- Do Experiment 21-A</u>

- Look at "Standard Reduction Table"

Ox agents on left + e⁻'s $\rightleftarrows$ Reducing agents on right

Stronger ox agents (More tendency to be reduced) (OAR) (To gain e⁻'s)	$F_2 + 2e^- \rightleftarrows 2F^-$	Stronger reducing agents (More tendency to be oxidized RAO) (To lose e⁻'s)
	$Ag^+ + e^- \rightleftarrows Ag\ (s)$	
	$Cu^{2+} + 2e^- \rightleftarrows Cu(s)$	
	$Zn^{2+} + 2e^- \rightleftarrows Zn\ (s)$	
	$Li^+ + e^- \rightleftarrows Li\ (s)$	

-So F_2 is a stronger ox agent than Ag+, etc.
-The strongest reducing agent on your chart is: _________.

Help in Hunting

- Solid metals mostly on bottom right (less active ones Ag, Au, farther up on the right side)
- Halogens (e.g. Cl_2) and oxyanions e.g. BrO_3^-, MnO_4^-, IO_3^- found near top left
- Some metal ions found on both sides e.g. Fe^{2+}, Sn^{2+}, Cu^+, Mn^{2+} can act as OA's or RA's

All the half-rx's are written as **reductions:**

e.g.) $F_2 + 2e^- \rightleftarrows 2F^-$
$Ag^+ + e^- \rightleftarrows Ag_{(s)}$

- The double arrow implies that **oxidation's** can also take place (reverse of reductions)

e.g.) reduction of Ag^+ (Same as table- single arrow)

Specifically just **reduction** → $Ag^+ + e^- \rightarrow Ag_{(s)}$

oxidation of Ag (**Reverse** of that on table- single arrow)

Specifically just **oxidation** → $Ag_{(s)} \rightarrow Ag^+ + e^-$

Write half-reactions for:
- Reduction of Pb^{2+} ________________________________
- Oxidation of Pb ________________________________
- Reduction of Sn^{2+} ________________________________
- Oxidation of Sn^{2+} ________________________________

- Oxidation of Fe^{2+} ______________________
- Reduction of Fe^{2+} ______________________
- Oxidation of Fe ______________________
- Reduction of acidified MnO_4- ______________________
- Oxidation of H_2 ______________________

Which is a stronger oxidizing agent: Ni^{2+} or Ag^{+}? ________
Fe^{2+} or Cr^{3+}? ________
Sn^{2+} or Sn^{4+}? ________

Must be on the left side when treating these as OA's

Which is a stronger reducing agent: Sn^{2+} or Fe^{2+}? ________
Zn or Ba? ________
Cl^{-} or Br^{-}? ________
Fe^{2+} or Au? ________

Must look for these on the right side

Which has a greater tendency to lose electrons, Ni or Zn? ______
Which has a greater tendency to gain electrons, Fe^{3+} or Cr^{3+}? ______
Which solid metal has the least tendency to lose electrons? ______
Which solid metal has the greatest tendency to lose e^{-}'s? ______
Give the formula for an ion that is a stronger oxidizing agent that Ni^{2+}, but is weaker than Pb^{2+}? ______

Using the reduction table to predict which reactions are spontaneous

- An **oxidizing agent** will react **spontaneously** with (oxidize) a reducing agent **<u>below</u> it on the <u>right</u>**

Look at your reduction chart!		
$F_2(g) + 2e^{-} \rightleftarrows 2F^{-}$	↓	F_2, the strongest OA, oxidize (react spontaneously with) **all** species below it on the right side from SO_4^{2-} all the way down to Li(s)
$S_2O_8 + 2e^{-} \rightleftarrows 2SO_4^{2-}$		
$Li^{+} + e^{-} \rightleftarrows Li_{(s)}$		

Look at the 4th half rx from the bottom		
$K^{+} + e^{-} \rightleftarrows K_{(s)}$	↓	K^{+} will oxidize only $Rb_{(s)}$, $Cs_{(s)}$ and $Li_{(s)}$, nothing else on the chart.
$Rb^{+} + e^{-} \rightleftarrows Rb_{(s)}$		
$Cs^{+} + e^{-} \rightleftarrows Cs_{(s)}$		
$Li^{+} + e^{-} \rightleftarrows Li_{(s)}$		

- A **reducing agent** on the right will react **spontaneously** with (reduce) any **oxidizing agent on the <u>left</u> <u>above</u> it**

e.g.) Li(s) (bottom right) will **reduce <u>all</u> species** on the **left** side except Li^{+}.
SO_4^{2-} (near top right) will reduce **only** F_2

- An OA on the left will **<u>not</u>** react spontaneously with a RA on the right above it!

e.g.) Au^{3+} will **<u>not</u>** oxidize (or react spontaneously with) SO_4^{2-}.

Some points…

1) Be very careful with charges e.g. Li^+ is a totally different thing than Li(s).

2) Things don't react with species which are **only on the same side** (these are **impossible** – not just non-spontaneous.)

   E.g.) K^+ (4th from bottom on the left) will **not** oxidize Rb^+ or Cs^+ Li^+ etc. –because they are on the **same side only**. (Impossible)

   E.g.) $Li_{(s)}$ will **not** reduce $Cs_{(s)}$, $Rb_{(s)}$, $K_{(s)}$, etc. because they are all on the **same side only**.

3) Some elements with **multiple oxidization** numbers e.g.) Sn, Cu, Mn, Fe have ions on **both sides** of the chart!
   –Look carefully at your table to find these.

> Note – Don't worry what E^o means yet, I will just use it to let you locate half-reactions.

Notice: Fe^{2+} is on the left (OA) at – 0.45
Fe^{2+} is on the right (RA) at + 0.77
Sn^{2+} is on the left (OA) at – 0.14
Sn^{2+} is on the right (RA) at + 0.15

A word about Cu…
Notice: Cu^+ is on the left at + 0.52
Cu^+ is on the right at + 0.15
-recall that anything on the left will oxidize a species below it on the right.

	E^o
$Cu^+ + e^- = Cu$	0.52
$Cu^{2+} + e^- = Cu^+$	0.15

This Cu^+ will oxidize this Cu^+

-Since Cu^+ **oxidizes** and **reduces** itself, any water solution of Cu^+ is **unstable** – it won't remain Cu^+ very long!

(demo Cu in HNO_3)

Notice: Mn^{2+} is on the left at $E^o = -1.19$
Mn^{2+} is on the right at $E^o = +1.22$

Also notice: $Cr^{3+} + e^- = Cr^{2+}$ - 0.41
$Cr^{3+} + 3e^- = Cr_{(s)}$ - 0.74

If a redox reaction is **non-spontaneous**, then the **reverse** reaction will be **spontaneous**!

e.g.) The reaction $Sr^{2+} + Ca_{(s)} \rightarrow Ca^{2+} + Sr_{(s)}$ is **non**-spontaneous because Ca is **above** Sr^{2+} on the **right** side.

But the rx: $Ca^{2+} + Sr_{(s)} \rightarrow Sr^{2+} + Ca_{(s)}$ **is** spontaneous because $Sr_{(s)}$ is **below** Ca^{2+} on the **right** side

Use the reduction table to answer the following questions:

a) Will Br_2 oxidize $Au_{(s)}$?………________________
b) Will $Pb_{(s)}$ reduce Fe^{2+}?……..________________
c) Will Zn^{2+} react with Cr^{3+}?…..________________
d) Will Mg^{2+} react with Cr^{3+}?…. ________________
e) Give the symbol of an ion that will oxidize $Mn_{(s)}$ but not $Cr_{(s)}$…..______________
f) Give the formula for a compound which will reduce Co^{2+} but will not reduce Fe^{2+}………………..________________
g) Which is a stronger reducing agent, Sn^{2+} or Fe^{2+}?(Hint – you must look for both on the ______ side)…..______________
h) Which is a stronger oxidizing agent, Cu^{+} or Sn^{2+}?(Hint – you must look for both on the _____ side)…..______________

Acidified solutions

-Any reactions on the table with $\mathbf{H^+}$ in them are **acidified** or **acid solutions**.

e.g.) Look at these: at E° = +1.51 (4th from the top)

$$\underbrace{MnO_4^- + 8H^+}_{\text{Called acidified permanganate solution}} + 5e^- \rightleftarrows Mn^{2+} + 4H_2O$$

Note: Names of many ions can be found on the ion table!

Give the E° corresponding to each of the following:

a) acidified iodate …………………..E°__________
b) acidified dichromate…………..….E°__________
c) acidified manganese (IV) oxide…E°__________
d) acidified bromate…………………E°__________
e) acidified perchlorate…………..….E°__________
f) acidified oxygen gas……………..E°__________

Nitric, Sulphuric & Phosphoric acids

- These acids are shown in **ionized form** on the table
- Nitric acid (HNO_3) is found in two places on the left side.

$NO_3^- + 4H^+ + 3e^- \rightleftarrows NO + 2H_2O$ $E^o = +0.96$ v

$2NO_3^- + 4H^+ + 2e^- \rightleftarrows N_2O_4 + 2H_2O$ $E^o = +0.80$ v

Don't worry about coefficients yet.
They are only used for balancing.

- Sulphuric acid is found at + 0.17 v

$SO_4^{2-} + 4H^+ + 2e^- \rightleftarrows H_2SO_3 + H_2O$ $E^o = +0.17$ v

Find and write the half-reaction for the reduction of **phosphoric** acid (H_3PO_4)

Sulphurous acid (H_2SO_3)

A note about water

$NO_3^- + 4H^+ + 3e^-$	$\rightleftarrows$	$NO_{(g)} + 2H_2O$	+0.96
$Hg^{2+} + 2e^-$	$\rightleftarrows$	$Hg_{(l)}$	+0.85
$\frac{1}{2}O_{2(g)} + 2H^+(10^{-7} M) + 2e^-$	$\rightleftarrows$	H_2O	+0.82
$2NO_3^- + 4H^+ + 2e^-$	$\rightleftarrows$	$N_2O_4 + 2H_2O$	+0.80
$Ag^+ + e^-$	$\rightleftarrows$	$Ag_{(s)}$	+0.80

-On the top of the table it says "ionic concentrations are at 1M"
-This includes $[H^+]$ = 1M with **two exceptions:**

-Neutral water is found on the **shaded lines** at + 0.82 v and – 0.41v
-Neutral water as a **reducing agent** is on the **right** side at + 0.82 v

-Neutral water as an **oxidizing agent** is on the **left** side at – 0.41 v

$Se_{(s)} + 2H^+ + 2e^-$	$\rightleftarrows$	H_2Se	−0.40
$Cr^{3+} + e^-$	$\rightleftarrows$	Cr^{2+}	−0.41
$2H_2O + 2e^-$	$\rightleftarrows$	$H_2 + 2OH^-(10^{-7} M)$	−0.41
$Fe^{2+} + 2e^-$	$\rightleftarrows$	$Fe_{(s)}$	−0.45

(Notice H_2O is below this at – 0.83 v but in this solution $[OH^-]$ = 1M (so it's basic, **not** neutral)

(Again H_2O is also found at + 1.23 v but here $[H^+]$ = 1M so it's acidic, **not** neutral)

Questions

a) Will neutral water oxidize Fe(s)? ________ Cr(s)? _______ Na(s)? _______
b) Will neutral water reduce Au^{3+}? _____ Ag^{+}? _____
c) Will acidified permanganate oxidize SO_4^{2-}? _____ Br^-? ______ Zn? _____
d) Will nitric acid react with Ag(s)? _____ Au(s)?_____ I^-_____Cl^-?______
e) Will nitric acid react with Fe^{2+}
f) Will nitric acid react with Hg to form N_2O_4? _____
g) Will nitric acid react with Hg to form NO? ______
h) Can you safely put a gold ring in acidified dichromate solution? _____ What about acidified bromate solution? _____
i) If Cl_2 gas is bubbled into water, will it all remain as Cl_2, or will some be converted to Cl^-? _____

Finding products of spontaneous reactions

eg) Given $Sn^{4+} + H_2S$ – find the products

See the table at +0.15v and +0.14v

$Sn^{4+} + 2e^- \rightleftarrows Sn^{2+}$ + 0.15v

$S(s) + 2H^+ + 2e^- \rightleftarrows H_2S$ + 0.14v

The **higher** reaction will be **reduction** (→), the lower reaction will proceed to the **left** (←) and be an **oxidation**.

$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$

$S(s) + 2H^+ + 2e^- \leftarrow H_2S$ (**reversed**! Lower one is **reversed**-is an **oxidation**)

-So the **products** are Sn^{2+}, S, and H^+

(at this point don't worry about coefficients yet.)

Questions

a) What are the products of the reaction of acidified hydrogen peroxide (H_2O_2) and bromide (Br^-)? ________________
b) What are the products of the reaction when neutral water reacts with:

Ca(s) ____________
Zn(s) ____________
Br_2 ______________
Acidified MnO_2 _____________
Fluorine gas _______________

Read SW p. 195-199
Do Ex 7-12 p. 199-200 SW