

Chemistry 12 – Unit 5

Oxidation – Reduction

Introduction

-Demonstration of oxidation – reduction reactions

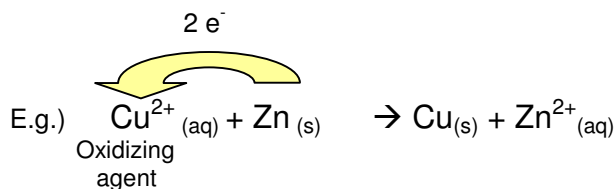
Definitions: (species means atom, ion or molecule)

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)



LEO says GER	
Losing Electrons is Oxidization	Gaining Electrons is Reduction

OAR
The <u>Oxidizing</u> <u>Agent</u> is <u>Reduced</u>

To carry it too far...

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

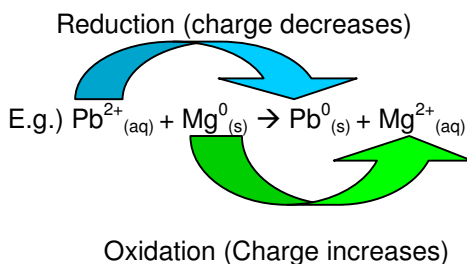
Redox – Short for **Oxidation – Reduction**

Redox identification

Charge on neutral atom or molecule = 0

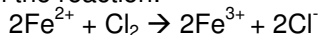
Oxidation – Charge gets more + (loses electrons)

Reduction – Charge gets more – (gains electrons)



Question

In the reaction:



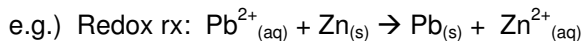
Identify:

- The Oxidizing Agent: _____
- The species being oxidized: _____
- The reducing agent: _____
- The species being reduced: _____
- The species gaining electrons: _____
- The species losing electrons: _____
- The product of oxidation _____
- 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.

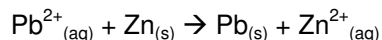


The Pb^{2+} (loses/gains) _____ 2 electrons.



Electrons on the LEFT side (or GER)
Means REDUCTION

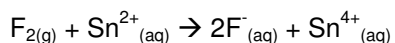
Write the **oxidation half reaction** for the following redox rx.



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:



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 **ionic compounds**

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

eg) $\underline{\text{Na}}\text{Cl}$ $\text{K}_2\underline{\text{CrO}}_4$

Ox # of Na & K = +1

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

eg) $\text{Ca}\underline{\text{Cl}}_2$ $\text{Al}\underline{\text{Br}}_3$ KF

Ox # of Cl, Br and F = -1

Note: Halogens are **not** always -1! (Only when it is written **last** in formula.)4) In **molecules or polyatomic ions**:

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

e.g.) $\text{K}\underline{\text{O}}\text{H}$ $\text{Cr}\underline{\text{O}}_4^{2-}$ $\text{Li}_3\underline{\text{P}}\underline{\text{O}}_4$ Ox # of
O is -2

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

Hydrogen Peroxide: $\text{H}_2\underline{\text{O}}_2$ **Alkali Peroxides:** $\text{Na}_2\underline{\text{O}}_2$

Ox # of O's = -1

(Remember, "O" in O_2 has an Ox. # of _____)5) In **molecules or ions**:a) **Hydrogen** is almost always +1e.g.) $\underline{\text{H}}\text{NO}_3$ $\underline{\text{H}}_2\underline{\text{S}}\text{O}_4$ $\underline{\text{H}}\underline{\text{P}}\underline{\text{O}}_4^{2-}$ } Every "H" has an ox # of +1b) An exception is **metallic hydrides**, which have an ox # of -1e.g.) $\underline{\text{Na}}\underline{\text{H}}$ $\text{Ca}\underline{\text{H}}_2$ (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 = _____)

Finding oxidation numbers of each atom in a molecule or PAI

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

e.g.) $\text{NH}_3 \leftarrow$ Total charge = 0 (no charge)

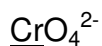
In a **polyatomic ion** – the **total ionic charge** is written on the **top right**

e.g.) $\text{CrO}_4^{2-} \leftarrow$ Total ionic charge (TIC) = -2

Oxidation numbers of all atoms add up to total ionic charge (TIC)

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

(Let x = ox # of one Cr atom)



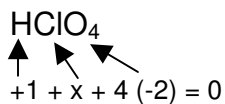
$$X + 4 [\text{\# of "O"atoms}] (-2 [\text{charge of oxygen}]) = -2 [\text{total ionic charge}]$$

$$X - 8 = -2$$

$$X = -2 + 8$$

$$X = \mathbf{+6}$$
 So ox # of Cr here = +6

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

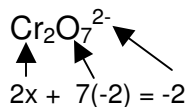


$$1 + x - 8 = 0$$

$$x - 7 = 0$$

$$x = \mathbf{+7}$$

e.g.) Find Ox # of Cr in $\text{Cr}_2\text{O}_7^{2-}$

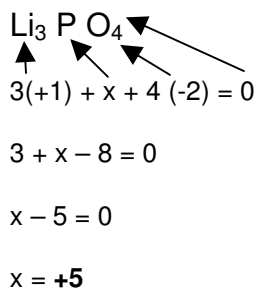


$$2x - 14 = -2$$

$$2x = +12$$

$$x = \mathbf{+6}$$

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



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

a) $\text{NaH}_2\underline{\text{P}}\text{O}_4$ _____ b) $\text{Na}_2\underline{\text{O}}_2$ _____ c) $\text{K}\underline{\text{H}}$ _____

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.

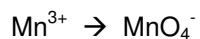
Changes in oxidation numbers

When an atom's **oxidation # is increased**, it is **oxidized**.

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

More complex:

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



- 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**.

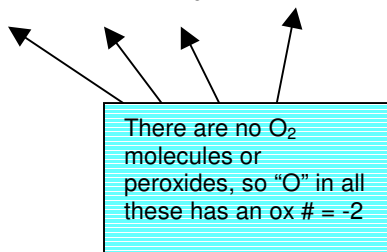
e.g.) $\text{Cu}(\text{NO}_3)_2 \rightarrow \text{Cu}_{(s)}$ Ox # **decreases** (reduction)

Ox # of Cu = +2	Ox # of Cu = 0
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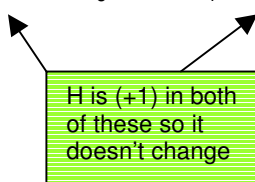
Redox ID using oxidation #'s

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

e.g.) $3\text{SO}_2 + 3\text{H}_2\text{O} + \text{ClO}_3^- \rightarrow 3\text{SO}_4^{2-} + 6\text{H}^+ + \text{Cl}^-$



$3\text{SO}_2 + 3\text{H}_2\text{O} + \text{ClO}_3^- \rightarrow 3\text{SO}_4^{2-} + 6\text{H}^+ + \text{Cl}^-$



Again:

$3\text{SO}_2 + 3\text{H}_2\text{O} + \text{ClO}_3^- \rightarrow 3\text{SO}_4^{2-} + 6\text{H}^+ + \text{Cl}^-$

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)

$3\text{SO}_2 \rightarrow 3\text{SO}_4^{2-}$

Coefficients are just for balancing.

SO_2

\rightarrow

SO_4^{2-}

Ox # of S is +4

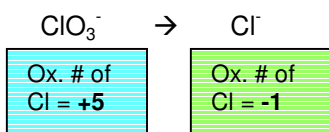
Ox # of S is +6

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

Note:

- R.A.O., the **reducing agent** is **oxidized**
- The species **SO₂** is acting as the **reducing agent**.
- The element **S** is being oxidized so **S is losing electrons**.

Look at the species with Cl:

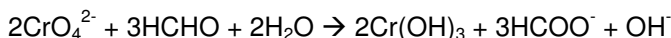


Decrease in ox # so Cl is being reduced

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^-)

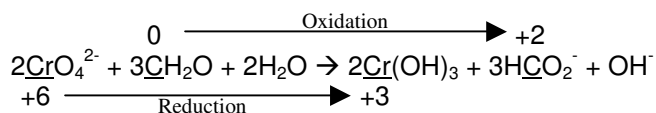
Eg. –given the reaction:



- 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

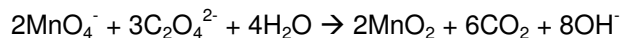
Notes:

- For hydrocarbons it's best to rewrite them as simple molecular formulas.
- All O's are in molecules or ions, no O₂ & no peroxides so O remains unchanged as -2
- All H's are in molecules or ions, no H₂ or metallic hydrides so H remains unchanged as +1
- The atoms to check for changes are C and Cr.



- So...
- a) the species being **oxidized** is (CH₂O) HCHO (inc. in ox #)
 - b) the **reducing agent** is (CH₂O) HCHO (RAO)
 - c) The species being **reduced** is CrO₄²⁻ (decrease in ox #)
 - d) The **oxidizing agent** is CrO₄²⁻ (OAR)
 - e) The species **losing e⁻s** is (CH₂O) HCHO (LEO)
 - f) The species **gaining e⁻s** is CrO₄²⁻ (GER)

Given the redox reaction:



Find:

- The species being reduced: _____.
- The species undergoing oxidation: _____.
- The oxidizing agent: _____.
- The reducing agent: _____.
- The species gaining electrons: _____.
- The species losing electrons: _____.

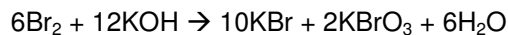
Given the balanced redox reaction:



Find:

- The oxidizing agent: _____.
- The reducing agent: _____.
- The species being reduced: _____.
- The species being oxidized: _____.
- The species losing electrons: _____.
- The species gaining electrons: _____.
- The product of oxidation: _____.
- The product of reduction: _____.

Given the following:



Find:

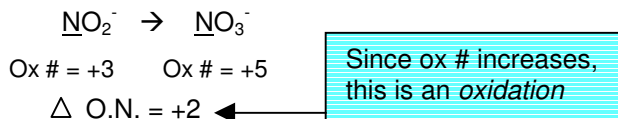
- The oxidizing agent: _____.
- The reducing agent: _____.
- The species undergoing oxidation: _____.
- The species being reduced: _____.
- The product of oxidation: _____.
- The product of reduction: _____.

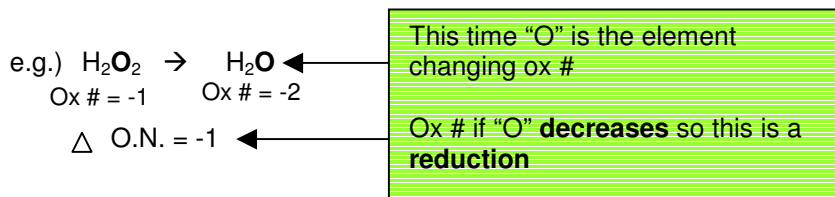
Using oxidation numbers to identify half-reactions

They don't have to be balanced

e.g.) If $\text{NO}_2^- \rightarrow \text{NO}_3^-$ 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.





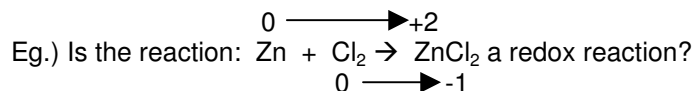
Find the $\Delta \text{O.N.}$ of the element in which it changes and identify each as an oxidation or reduction

- a) $\text{C}_2\text{H}_5\text{OH} \rightarrow \text{CH}_3\text{COOH}$ _____
 b) $\text{Fe}_2\text{O}_3 \rightarrow \text{Fe}_3\text{O}_4$ _____
 c) $\text{H}_3\text{PO}_4 \rightarrow \text{P}_4$ _____
 (P_4 is the elemental form of phosphorus)
 d) $\text{CH}_3\text{COOH} \rightarrow \text{CH}_3\text{COH}$ _____

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 **always** be redox, because in **elemental form ox. # = 0** and in compounds usually **ox. # is not = 0**



Answer: It must be because ΔON of Zn ($0 \rightarrow +2 = +2$) and ΔON of Cl ($0 \rightarrow -1 = -1$)

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

Half-reactions and the reduction table

- Do Experiment 21-A

- Look at “Standard Reduction Table”

Ox agents on left + e ⁻ s ⇌ Reducing agents on right			
Stronger ox agents (More tendency to be reduced) (OAR) (To gain e ⁻ s)	↑	$F_2 + 2e^- \rightleftharpoons 2F^-$	Stronger reducing agents (More tendency to be oxidized RAO) (To lose e ⁻ s)
		$Ag^+ + e^- \rightleftharpoons Ag(s)$	
		$Cu^{2+} + 2e^- \rightleftharpoons Cu(s)$	
		$Zn^{2+} + 2e^- \rightleftharpoons Zn(s)$	
		$Li^+ + e^- \rightleftharpoons Li(s)$	

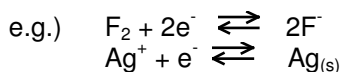
-So F₂ is a stronger ox agent than Ag⁺, etc.

-The strongest reducing agent on your chart is: _____.

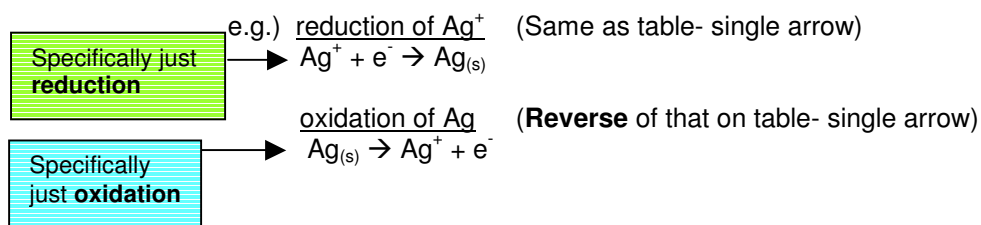
Help in Hunting

- Solid metals mostly on bottom right (less active ones Ag, Au, farther up on the right side)
- Halogens (e.g. Cl₂) and oxyanions e.g. BrO₃⁻, MnO₄⁻, IO₃⁻ found near top left
- Some metal ions found on both sides e.g. Fe²⁺, Sn²⁺, Cu⁺, Mn²⁺ can act as OA's or RA's

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



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



Write half-reactions for:

- Reduction of Pb²⁺
- Oxidation of Pb
- Reduction of Sn²⁺
- Oxidation of Sn²⁺

- 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 than 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 **below it on the right**

Look at your reduction chart!	
$\text{F}_2(\text{g}) + 2\text{e}^- \rightleftharpoons 2\text{F}^-$	↓ 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)
$\text{S}_2\text{O}_8 + 2\text{e}^- \rightleftharpoons 2\text{SO}_4^{2-}$	
$\text{Li}^+ + \text{e}^- \rightleftharpoons \text{Li}(\text{s})$	

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

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

e.g.) Li(s) (bottom right) will **reduce all species** on the **left** side except Li⁺.
 SO_4^{2-} (near top right) will reduce **only** F₂

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

e.g.) Au^{3+} will **not** 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 oxidation** 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° means yet, I will just use it to let you locate half-reactions.

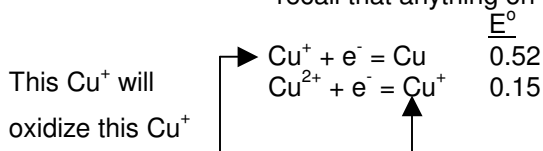
Notice: Fe^{2+} is on the left (OA) at -0.45
 Fe^{2+} is on the right (RA) at $+0.77$
 Sn^{2+} is on the left (OA) at -0.14
 Sn^{2+} is on the right (RA) at $+0.15$

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.



-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 $E^\circ = -1.19$
 Mn^{2+} is on the right at $E^\circ = +1.22$

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

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

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

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

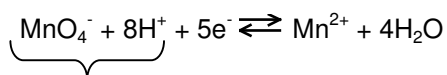
Use the reduction table to answer the following questions:

- Will Br_2 oxidize $\text{Au}_{(s)}$?.....
- Will $\text{Pb}_{(s)}$ reduce Fe^{2+} ?.....
- Will Zn^{2+} react with Cr^{3+} ?.....
- Will Mg^{2+} react with Cr^{3+} ?.....
- Give the symbol of an ion that will oxidize $\text{Mn}_{(s)}$ but not $\text{Cr}_{(s)}$
- Give the formula for a compound which will reduce Co^{2+} but will not reduce Fe^{2+}
- Which is a stronger reducing agent, Sn^{2+} or Fe^{2+} ? (Hint – you must look for both on the _____ side).....
- 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 H^+ in them are **acidified** or **acid solutions**.

e.g.) Look at these: at $E^\circ = +1.51$ (4th from the top)



Called **acidified** permanganate solution

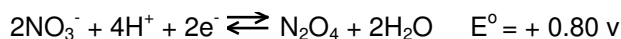
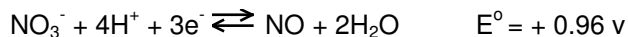
Note: Names of many ions can be found on the ion table!

Give the E° corresponding to each of the following:

- acidified iodate E° _____
- acidified dichromate..... E° _____
- acidified manganese (IV) oxide... E° _____
- acidified bromate..... E° _____
- acidified perchlorate..... E° _____
- acidified oxygen gas..... E° _____

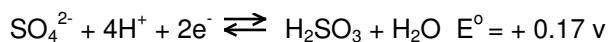
Nitric, Sulphuric & Phosphoric acids

- These acids are shown in **ionized form** on the table
- Nitric acid (HNO₃) is found in two places on the left side.



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

- Sulphuric acid is found at + 0.17 v



Find and write the half-reaction for the reduction of **phosphoric acid** (H₃PO₄)

Sulphurous acid (H₂SO₃)

A note about water

-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

$\text{NO}_3^- + 4\text{H}^+ + 3\text{e}^- \rightleftharpoons \text{NO}(\text{g}) + 2\text{H}_2\text{O}$	+0.96
$\text{Hg}^{2+} + 2\text{e}^- \rightleftharpoons \text{Hg}(\text{l})$	+0.85
$\frac{1}{2}\text{O}_2(\text{g}) + 2\text{H}^+(10^{-7}\text{M}) + 2\text{e}^- \rightleftharpoons \text{H}_2\text{O}$	+0.82
$2\text{NO}_3^- + 4\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{N}_2\text{O}_4 + 2\text{H}_2\text{O}$	+0.80
$\text{Ag}^+ + \text{e}^- \rightleftharpoons \text{Ag}(\text{s})$	+0.80

$\text{Se}(\text{l}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{Se}$	-0.40
$\text{Cr}^{3+} + \text{e}^- \rightleftharpoons \text{Cr}^{2+}$	-0.41
$2\text{H}_2\text{O} + 2\text{e}^- \rightleftharpoons \text{H}_2 + 2\text{OH}^-(10^{-7}\text{M})$	-0.41
$\text{Fe}^{2+} + 2\text{e}^- \rightleftharpoons \text{Fe}(\text{l})$	-0.45

(Notice H₂O is below this at - 0.83 v but in this solution [OH⁻] = 1M (so it's **basic**, **not** neutral)

(Again H₂O 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 Au³⁺? _____ Ag⁺? _____
 c) Will acidified permanganate oxidize SO₄²⁻? _____ Br⁻? _____ Zn? _____
 d) Will nitric acid react with Ag(s)? _____ Au(s)? _____ I⁻? _____ Cl⁻? _____
 e) Will nitric acid react with Fe²⁺? _____
 f) Will nitric acid react with Hg to form N₂O₄? _____
 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 Cl₂ gas is bubbled into water, will it all remain as Cl₂, or will some be converted to Cl⁻? _____

Finding products of spontaneous reactions

- eg) Given Sn⁴⁺ + H₂S – find the products
 See the table at +0.15v and +0.14v
 $\text{Sn}^{4+} + 2\text{e}^- \rightleftharpoons \text{Sn}^{2+} + 0.15\text{v}$
 $\text{S}(\text{s}) + 2\text{H}^+ + 2\text{e}^- \rightleftharpoons \text{H}_2\text{S} + 0.14\text{v}$
 The **higher** reaction will be **reduction** (→), the lower reaction will proceed to the **left** (←) and be an **oxidation**.
 $\text{Sn}^{4+} + 2\text{e}^- \rightarrow \text{Sn}^{2+}$
 $\text{S}(\text{s}) + 2\text{H}^+ + 2\text{e}^- \leftarrow \text{H}_2\text{S}$ (**reversed!** Lower one is **reversed**-is an **oxidation**)
 -So the **products** are Sn²⁺, 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 (H₂O₂) and bromide (Br⁻)? _____
 b) What are the products of the reaction when neutral water reacts with:
 Ca(s) _____
 Zn(s) _____
 Br₂ _____
 Acidified MnO₂ _____
 Fluorine gas _____

Read SW p. 195-199

Do Ex 7-12 p, 199-200 SW