Using data to make your own simple Redox table

Example problem:

1) Four metals A, B, C, & D were tested with separate solutions of A^{2+}, B^{2+}, C^{2+} & D^{2+}. Some of the results are summarized in the following table:

	Solution			
Metal	A^{2+}	B^{2+}	C^{2+}	D^{2+}
A		(1) no reaction	(2) reaction	
B				(4) no reaction
D	(3) reaction			

List the ions in order from the strongest to weakest oxidizing agent.

NOTE: For the same element: The more positive species is always the Oxidizing Agent.
Eg.) A^{2+} A
OA RA

Using data

1) – Since B^{2+} does not oxidize A : B^{2+} must be below A on the table.

Oxidizing agents | Reducing agents

$$A^{2+} + 2e^- = A$$
$$B^{2+} + 2e^- = B$$

2) – Since C^{2+} reacts with A: C^{2+} must be above A:

$$C^{2+} + 2e^- = C$$
$$A^{2+} + 2e^- = A$$
$$B^{2+} + 2e^- = B$$

3) – Since A^{2+} reacts with D: A^{2+} must be above D on the table. But is D^{2+} above or below B^{2+}? We don't know yet.

C^{2+}
D^{2+} Here → A^{2+}
B^{2+}
Or here?

Let's look at the next information:

4) – D^{2+} does not react with B

- Now we know that D^{2+} must be below B on the table

So now we have our complete table:

Oxidizing agents	Reducing agents
$C^{2+} + 2e^- = C$	
$A^{2+} + 2e^- = A$	
$B^{2+} + 2e^- = B$	
$D^{2+} + 2e^- = D$	

- At this point its good to go back and recheck that all the data given is consistent with your table.
- So now we have our answer; The ions in order of strongest to weakest ox agent is: C^{2+}, A^{2+}, B^{2+}, D^{2+}
- Just in case you're asked, you can see that the order of reducing agent from strongest to weakest is D, B, A, C.

<u>Another example –</u>

Four non-metal oxidizing agents X_2, Y_2, Z_2 and W_2 are combined with solutions of ions: X^-, Y^-, Z^- and W^-.

NOTE: For the same element: The more positive species is always the Oxidizing Agent.
Eg.) X_2 (OA) X^- (RA)

The following results were obtained;

(1) X_2 reacts with W^- and Y^- only.
(2) Y^- will reduce W_2

List the reducing agents from strongest to weakest

(1) X_2 will be above W^- & Y^-, but below Z^-

Oxidizing agents		Reducing agents
$Z_2 + 2e^-$	$\rightleftarrows$	$2Z^-$
$X_2 + 2e^-$	$\rightleftarrows$	$2X^-$

W^- & Y^- } Are both below X_2, but we don't know in which order yet.

(2) Since Y^- reduces W_2, Y^- must be lower on the right of W_2.

OA's		RA's
$Z_2 + 2e^-$	$\rightleftarrows$	$2Z^-$
$X_2 + 2e^-$	$\rightleftarrows$	$2X^-$
$W_2 + 2e^-$	$\rightleftarrows$	$2W^-$
$Y_2 + 2e^-$	$\rightleftarrows$	$2Y^-$

To answer the question:

The <u>reducing agents</u> from strongest to weakest are: <u>Y^-, W^-, X^-, Z^-</u>

<u>Question:</u>

Four solutions $A(NO_3)_2$, $B(NO_3)_2$, $C(NO_3)_2$, and $D(NO_3)_2$ are added to metals, A, B, C, & D

The following information is found:

(1) The metal A will not react with any of the solutions
(2) $C(NO_3)_2$ reacts spontaneously with B
(3) B will not react with $D(NO_3)_2$

(a) Make a small reduction table showing reductions of the metallic ions. (Don't forget to **discard** the **spectator** nitrate ions.

(b) List the oxidizing agents in order of strongest to weakest: ____________________

(c) List the reducing agent in order of strongest to weakest: ____________________

(d) Would it be safe to store $A(NO_3)_2$ solution in a container made of the metal D? _______________

Do Exercises 14,15,16 & 18 on p. 200 of SW.

Balancing half-reactions

-Some half-rx's are on the table, but not all.
-Given if the soln. Is acidic or basic.

Pay attention!

-Think of ***Major Hydroxide*** (Major → O → H → - (charge))

Major atoms →atoms other than O & H

Acid Soln. E.g.) $S_2O_8^{2-} \rightarrow HSO_3^-$ (acid soln.)

(1) Balance Major Atoms (S in this case)

$S_2O_8^{2-} \rightarrow 2HSO_3^-$

(2) Balance "O" atoms, by adding H_2O (to the side with less O's)

$S_2O_8^{2-} \rightarrow 2HSO_3^- + 2H_2O$

(3) Balance "H" atoms by adding H^+ (to the side with less H's)

$S_2O_8^{2-} + 6H^+ \rightarrow 2HSO_3^- + 2H_2O$

(4) Balance charge by adding e^-'s (to the more + side)

$S_2O_8^{2-} + 6H^+ \rightarrow 2HSO_3^- + 2H_2O$

TIC = (+4) TIC = (-2)

The left side needs $6e^-$'s to get a –2 charge

So the final balanced half-rx is:

$S_2O_8^{2-} + 6H^+ + 6e^- \rightarrow 2HSO_3^- + 2H_2O$

TIC = (-2) TIC = (-2)

-Always double-check these!
-Don't miscopy charges, etc.

Try this one: $MnO_4^- \rightarrow Mn^{2+}$ (acid soln)

In basic solution

-Do the first steps of the balancing just like an acid

E.g.) $MnO_2 \rightarrow MnO_4^-$ (basic solution)

Major (Mn already balanced)
Oxygen $2H_2O + MnO_2 \rightarrow MnO_4^-$
Hydrogen $2H_2O + MnO_2 \rightarrow MnO_4^- + 4H^+$
Charge $2H_2O + MnO_2 \rightarrow MnO_4^- + 4H^+ + 3e^-$

In basic solution: write the reaction $H^+ + OH^- \rightarrow H_2O$ or $H_2O \rightarrow H^+ + OH^-$

-In whichever way is needed to cancel out the H^+'s
-Add to the half-rx

E.g.) $2H_2O + MnO_2 \rightarrow MnO_4^- + 4H^+ + 3e^-$
$4H^+ + 4OH^- \rightarrow 4H_2O$

We need $4H^+$ on the left to cancel the $4H^+$ on the left side

You must write the whole water equation

(Take away $2H_2O$'s from both sides)

$MnO_2 + 4OH^- \rightarrow MnO_4^- + 2H_2O + 3e^-$

Double-check everything!

Try this one: $Pb \rightarrow HPbO_2^-$ (basic soln)

-Reactions without H's or O's are done in neutral soln -Do Ex 19 a-m p. 203

Balancing overall redox reactions using the half-reaction (half-cell) method

(1) Break up Rx into 2 half-rx's.
(2) Balance each one (in acidic or basic as given)
(3) Multiply each half rx by whatever is needed to cancel out e^-'s
(Note: balanced half-rx have e^-'s (on left reduction on right oxidation) Balanced redox don't have e^-'s)
(4) Add the 2 half-rx's canceling e^-'s and anything else (usually H_2O's, H^+'s or OH^-'s) in order to simplify.

Example: $U^{4+} + MnO_4^- \rightarrow Mn^{2+} + UO_2^{2+}$ (acidic)

Balance each ½ rx

$U^{4+} \rightarrow UO_2^{2+}$
(Major (U) balanced already)
Oxygen → $U^{4+} + 2H_2O \rightarrow UO_2^{2+}$
Hydrogen → $U^{4+} + 2H_2O \rightarrow UO_2^{2+} + 4H^+$
Charge → $U^{4+} + 2H_2O \rightarrow UO_2^{2+} + 4H^+ + 2e^-$

$MnO_4^- \rightarrow Mn^{2+}$
(Major (Mn) balanced already)
$MnO_4^- \rightarrow Mn^{2+} + 4H_2O$
$MnO_4^- + 8H^+ \rightarrow Mn^{2+} + 4H_2O$
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$

$(U^{4+} + 2H_2O \rightarrow UO_2^{2+}\ 4H^+ + 2e^-)\ 5$ ← Multiply by 5 to get $10e^-$
$(MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O)\ 2$ ← Multiply by 2 to get $10e^-$

$$5U^{4+} + 10H_2O + 2MnO_4^- + 16H^+ + 10e^- \rightarrow 5UO_2^{2+} + 20H^+ + 2Mn^{2+} + 8H_2O + 10e^-$$

To simplify:
-Take away $10e^-$ from both sides
-Take away $16H^+$'s from both sides
-Take away $8H_2O$'s from both sides

$$5U^{4+} + 2H_2O + 2MnO_4^- \rightarrow 5UO_2^{2+} + 4H^+ + 2Mn^{2+}$$

Quick check by finding TIC's on both sides

+20 + 0 – 2   +10 + 4 + 4

TIC = +18   TIC = +18

- If you have time check all atoms also if TIC's are not equal you messed up! Somewhere! Find it!

Try this one:

$SO_2 + IO_3^- \rightarrow SO_4^{2-} + I_2$ (basic solution)

-See examples p.205-207 in SW

<u>Quick notes</u>

-Some redox equations have just **one reactant**

- Use this as the reactant in **both** half-rx's.
- These are called "self-oxidation-reduction" or Disproportionation reactions.

Eg) $Br_2 \rightarrow Br^{--} + BrO_3^{-}$ (basic) (found in some hot tubs)

<u>Half rx's are:</u>

$Br_2 \rightarrow Br^{--}$	$Br_2 \rightarrow BrO_3^{-}$

Answer: ____

Do Ex 24 a-w p. 207
<u>The more practice the better!</u> <u>See me if you want more!</u>

Balancing redox equations using the oxidation number method

-<u>This is optional</u>

- As long as one method (not guessing!) works for you that's fine. (This method or half-rx method.)
- Read examples p. 271-272 SW
- Do any ex 10 a-n & check with key

Redox titrations

- same as in other units (solubility/acids-bases)
- coefficient ratios for the "mole bridge" are obtained by the balanced redox equation:

TITRATIONS		
STANDARD	mole bridge	*SAMPLE*
Conc. & Volume → *moles* or Mass → *moles*	→	*moles* → Conc. or Volume
mol = M x L or: grams x $\frac{1 \text{ mol}}{\text{MM g}}$ = mol		M = mol/L or L = mol/M

Eg) Acidified hydrogen peroxide (H_2O_2) is used to titrate a solution of MnO_4^- ions of unknown concentration. Two products are O_2 gas and Mn^{2+}.

a) Write the **balanced redox equation**:

b) It takes 6.50 mL of 0.200 M H_2O_2 to titrate a 25.0 mL sample of MnO_4^- solution. Calculate the original $[MnO_4^-]$.

Finding a suitable solution titrate a sample

Use redox table:

- If sample is on the **left** (OA)
  Use something **below it on the right**. (RA)
- If sample is on the **right** (RA) use something **above it on the left** (OA)
- Good standards will **change colour** as they react

Acidified MnO_4^- (purple) = Mn^{2+} (clear)
Acidified $Cr_2O_7^{2-}$ (orange) = Cr^{3+} (pale green)

Read p. 210-212 carefully – go over the examples! Do ex 26 & 29 p. 213-214 SW.