

## Using data to make your own simple Redox table

Example problem:

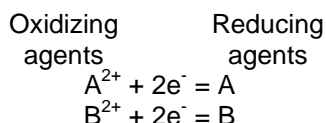
- 1) Four metals A, B, C, & D were tested with separate solutions of  $A^{2+}$ ,  $B^{2+}$ ,  $C^{2+}$  &  $D^{2+}$ .  
Some of the results are summarized in the following table:

| Metal | Solution                |                            |                         |                            |
|-------|-------------------------|----------------------------|-------------------------|----------------------------|
|       | $A^{2+}$                | $B^{2+}$                   | $C^{2+}$                | $D^{2+}$                   |
| A     |                         | <sup>(1)</sup> no reaction | <sup>(2)</sup> reaction |                            |
| B     |                         |                            |                         | <sup>(4)</sup> no reaction |
| D     | <sup>(3)</sup> reaction |                            |                         |                            |

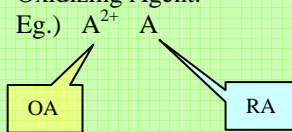
List the ions in order from the strongest to weakest oxidizing agent.

Using data

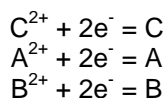
- 1) – Since  $B^{2+}$  does not oxidize A :  $B^{2+}$  must be below A on the table.



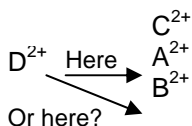
NOTE: For the same element: The more positive species is always the Oxidizing Agent.



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



- 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.



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.

### Another example –

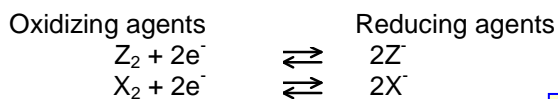
Four non-metal oxidizing agents  $X_2$ ,  $Y_2$ ,  $Z_2$  and  $W_2$  are combined with solutions of ions:  $X^-$ ,  $Y^-$ ,  $Z^-$  and  $W^-$ .

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  $Z^-$



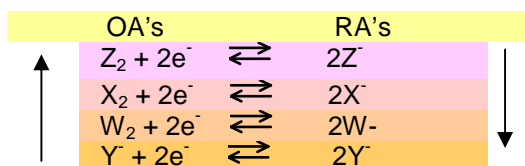
$W^-$  &  $Y^-$  } Are both below  $X_2$ ,  
but we don't know in  
which order yet.

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

OA

RA

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



To answer the question:

The reducing agents from strongest to weakest are:  $Y^-$ ,  $W^-$ ,  $X^-$ ,  $Z^-$

### Question:

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  $\text{A}(\text{NO}_3)_2$  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**.

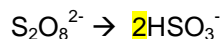
Pay attention!

-Think of **Major Hydroxide** (Major  $\rightarrow \text{O} \rightarrow \text{H} \rightarrow -$  (charge))

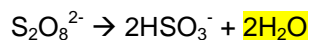
**Major atoms  $\rightarrow$  atoms other than O & H**

Acid Soln. E.g.)  $\text{S}_2\text{O}_8^{2-} \rightarrow \text{HSO}_3^-$  (acid soln.)

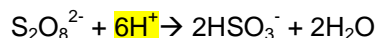
(1) Balance Major Atoms (S in this case)



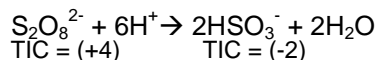
(2) Balance "O" atoms by adding  $\text{H}_2\text{O}$  (to the side with less O's)



(3) Balance "H" atoms by adding  $\text{H}^+$  (to the side with less H's)

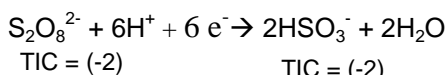


(4) Balance charge by adding  $\text{e}^-$ 's (to the more + side)



The left side needs  $6\text{e}^-$ 's  
to get a -2 charge

**So the final balanced half-rx is:**



-Always double-check these!

-Don't miscopy charges, etc.

Try this one:  $\text{MnO}_4^- \rightarrow \text{Mn}^{2+}$  (acid soln)

### In basic solution

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

E.g.)  $\text{MnO}_2 \rightarrow \text{MnO}_4^-$  (basic solution)

Major (Mn already balanced)

Oxygen  $2\text{H}_2\text{O} + \text{MnO}_2 \rightarrow \text{MnO}_4^-$

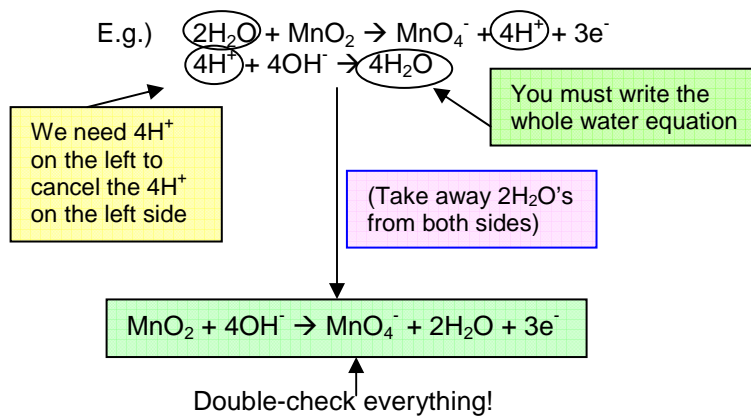
Hydrogen  $2\text{H}_2\text{O} + \text{MnO}_2 \rightarrow \text{MnO}_4^- + 4\text{H}^+$

Charge  $2\text{H}_2\text{O} + \text{MnO}_2 \rightarrow \text{MnO}_4^- + 4\text{H}^+ + 3\text{e}^-$

In basic solution: write the reaction  $\text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O}$  or  $\text{H}_2\text{O} \rightarrow \text{H}^+ + \text{OH}^-$

-In whichever way is needed to cancel out the  $\text{H}^+$ 's

-Add to the half-rx



Try this one:  $\text{Pb} \rightarrow \text{HPbO}_2^-$  (basic soln)

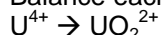
-Reactions without H's or O's are done in **neutral** soln -Do Ex 19 a-m p. 203

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

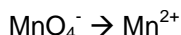
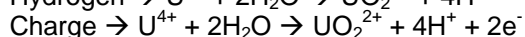
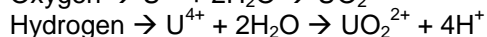
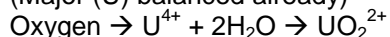
- (1) Break up Rx into 2 half-rx's.
- (2) Balance each one (in acidic or basic as given)
- (3) Multiply each half rx by whatever is needed to cancel out  $e^-$ 's  
(Note: balanced half-rx have  $e^-$ 's (on left reduction on right oxidation) Balanced redox don't have  $e^-$ 's)
- (4) Add the 2 half-rx's 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)

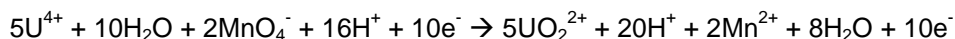
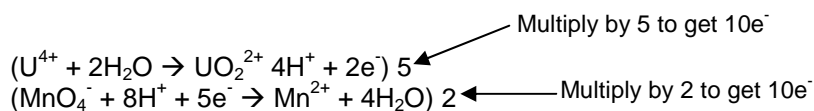
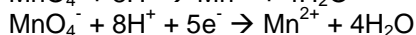
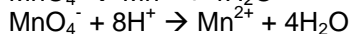
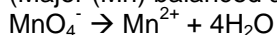
Balance each  $\frac{1}{2}$  rx



(Major (U) balanced already)

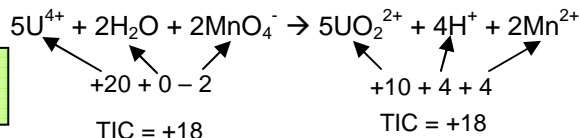


(Major (Mn) balanced already)



To simplify:

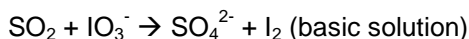
- Take away  $10e^-$  from both sides
- Take away  $16H^+$ 's from both sides
- Take away  $8H_2O$ 's from both sides



Quick check by finding  
TIC's on both sides

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

Try this one:



-See examples p.205-207 in SW

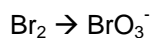
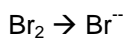
### Quick notes

-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)  $\text{Br}_2 \rightarrow \text{Br}^- + \text{BrO}_3^-$  (basic) (found in some hot tubs)

Half rx's are:



Answer: \_\_\_\_\_

Do Ex 24 a-w p. 207

The more practice the better! See me if you want more!

### Balancing redox equations using the oxidation number method

-This is optional

- 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   | SAMPLE                                      |
| Conc. & Volume $\rightarrow$ moles<br>or Mass  | moles $\rightarrow$ Conc. or Volume         |
| mole bridge  |   |
| $\text{mol} = M \times L$<br>or: grams $\times \frac{1 \text{ mol}}{\text{MM g}} = \text{mol}$ | $M = \text{mol/L}$<br>or $L = \text{mol/M}$ |

Eg) Acidified hydrogen peroxide ( $\text{H}_2\text{O}_2$ ) is used to titrate a solution of  $\text{MnO}_4^-$  ions of unknown concentration. Two products are  $\text{O}_2$  gas and  $\text{Mn}^{2+}$ .

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

b) It takes 6.50 mL of 0.200 M  $\text{H}_2\text{O}_2$  to titrate a 25.0 mL sample of  $\text{MnO}_4^-$  solution. Calculate the original  $[\text{MnO}_4^-]$ .

### 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  $\text{MnO}_4^-$  (purple) =  $\text{Mn}^{2+}$  (clear)

Acidified  $\text{Cr}_2\text{O}_7^{2-}$  (orange) =  $\text{Cr}^{3+}$  (pale green)

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