

Additional  
EC & Kinetics

Q's

for review purposes

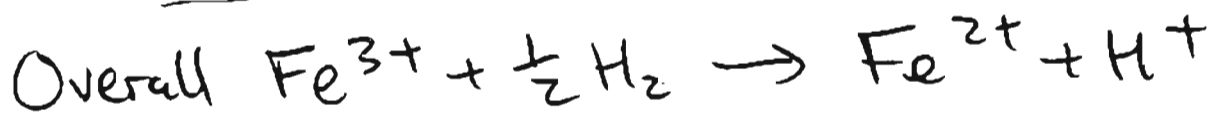
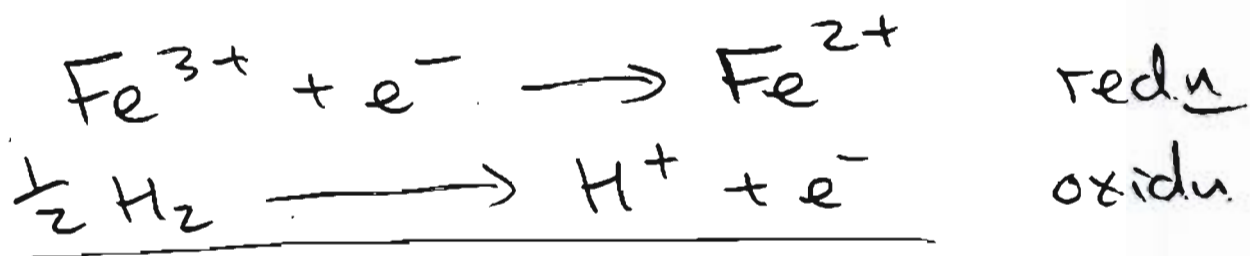


EC/Kin ©

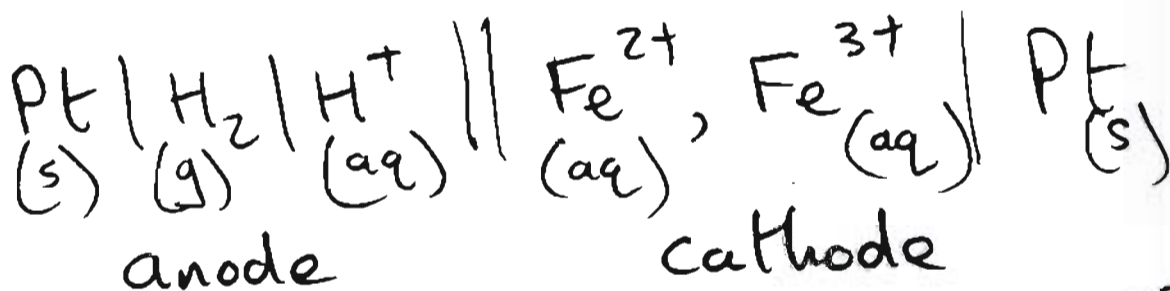
## ADDITIONAL EC Q's.

Q Write the cell diagram for the rx in which hydrogen is used to reduce  $\text{Fe}^{3+}(\text{aq})$  to  $\text{Fe}^{2+}(\text{aq})$

Strategy - write overall cell rx. and/or break into oxidn and redn half reactions. [Note - if a redox couple does not include a solid metal use Pt for electrical contact]



∴ cell is



EC(1)

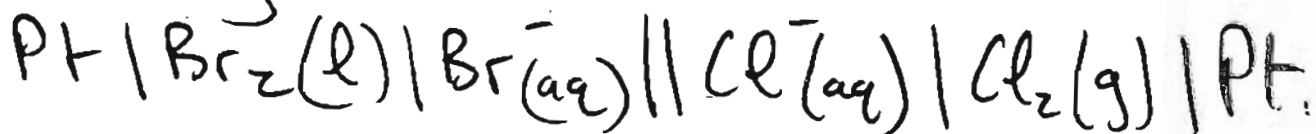
Q Write the cell diag. for a cell in which one electrode is  $\text{Pt} | \text{Cl}_2(\text{g}) | \text{Cl}^-(\text{aq})$  and the other  $\text{Pt} | \text{Br}_2(\text{l}) | \text{Br}^-$ . Give the cell rx.

Strategy

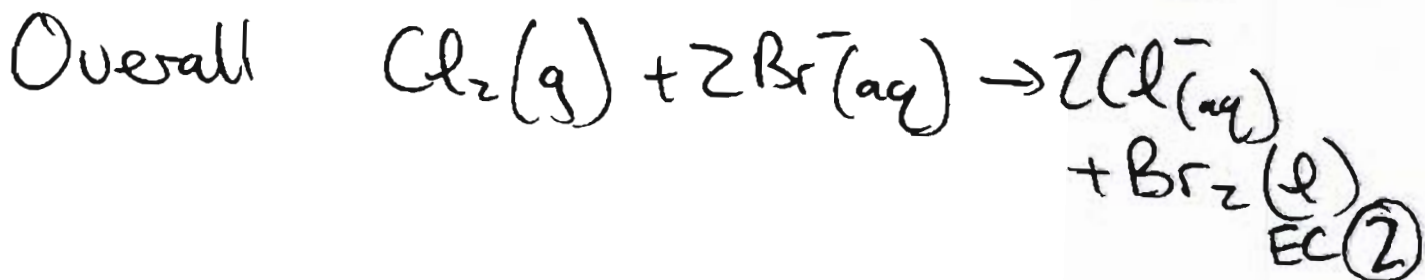
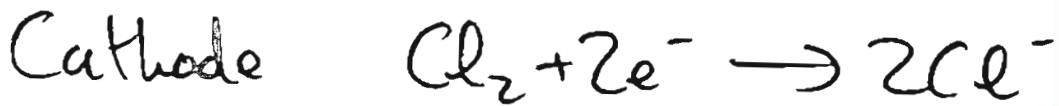
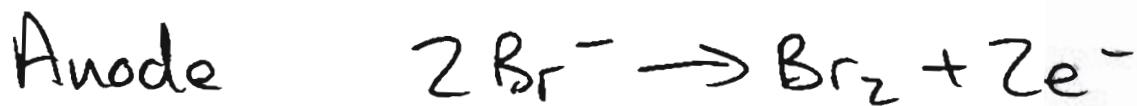
$\text{Cl}_2$  is more <sup>than</sup> +ve  $\text{Br}_2$  in the SRP Table. i.e. better oxidising agent.

$\therefore \text{Cl}_2$  will be the cathode ('cos it will be reduced)

Cell diag is:

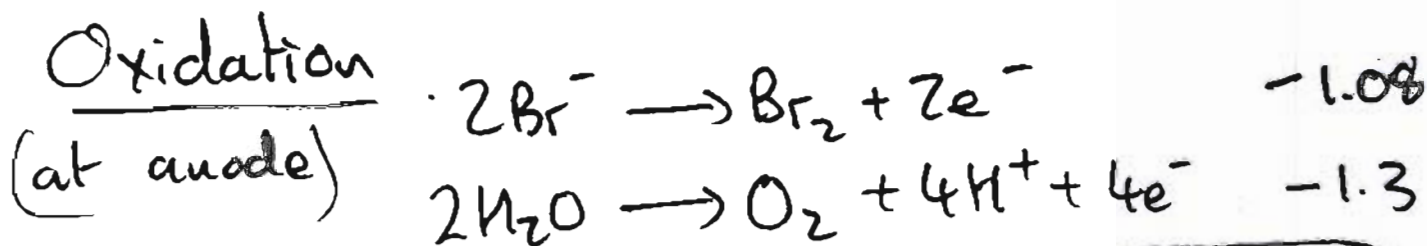
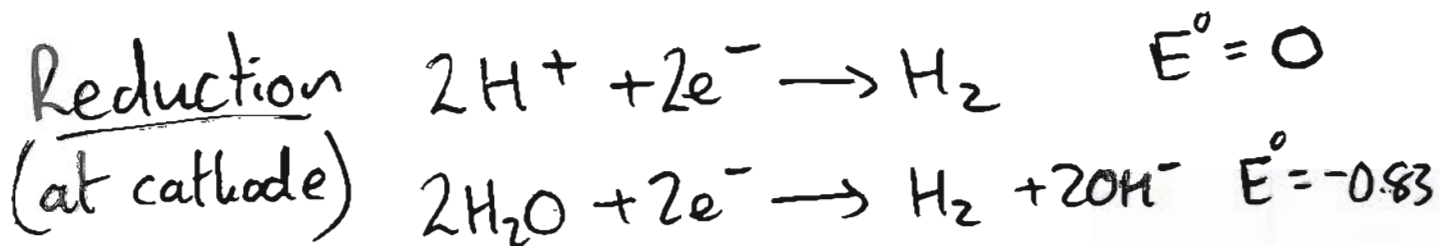


Half rxs are



Q Write balanced net ionic equations for the anode and cathode reactions in the electrolysis of HBr(aq)

Strategy - identify species in solution, and write possible oxidation and reduction reactions.



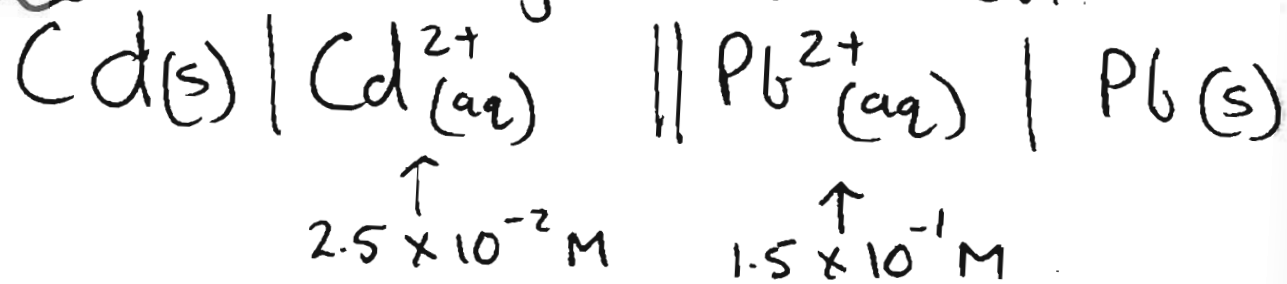
Look at Table: OR, no-brainer... the most +ve in both cases, wins more easily reduced

For reduction:  $\text{H}^+$  is a better oxid. agent (more easily reduced)  
 $\therefore$  first rx. wins.

For oxidation:  $\text{Br}^-$  is a better red. agent (more easily oxid.)  
 $\therefore$  first equation wins.

EC(3)

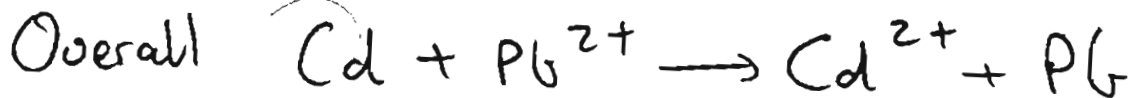
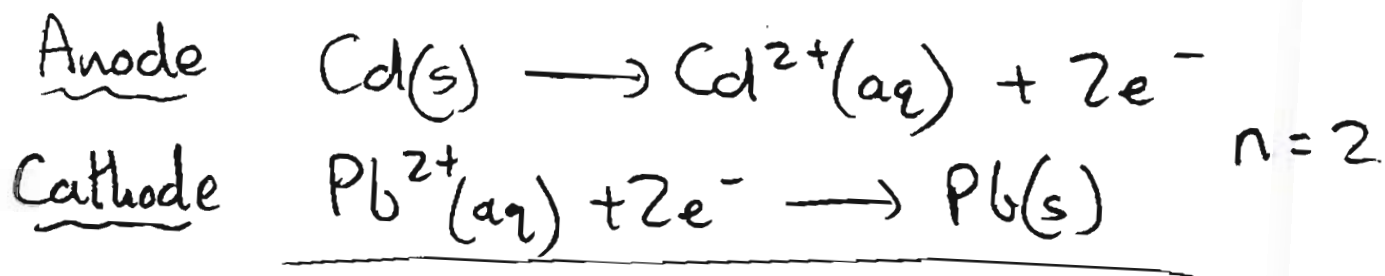
Q The voltage of the cell



is 0.293 V at 298 K

---

(a) Calculate  $\Delta G$  per mol of Pb produced at these concns



$$\begin{aligned} \Delta G &= -nFE = -(2 \text{ mol})(96,485 \text{ J V}^{-1} \text{ mol}^{-1})(0.293 \text{ V}) \\ &= \underline{\underline{-5.65 \times 10^4 \text{ J}}} \end{aligned}$$

(b) Calculate  $\Delta G^\circ$  per mol of Pb produced

$$E = E^\circ - \frac{RT}{nF} \ln Q$$

$$\text{or } E = E^\circ - \frac{0.0592}{n} \log Q \quad ; \quad Q = \frac{[\text{Cd}^{2+}]}{[\text{Pb}^{2+}]}$$

$$0.293 = E^\circ - \frac{0.0592}{2} \log \frac{2.5 \times 10^{-2}}{1.5 \times 10^{-1}}$$

$$\therefore E^\circ = 0.270 \text{ V}$$

$$\begin{aligned} \Delta G^\circ &= -nFE^\circ = -(2)(96,485)(0.270) \\ &= \underline{\underline{-5.21 \times 10^4 \text{ J}}} \end{aligned}$$

Note - could have used

$\Delta G = \Delta G^\circ + RT \ln Q$  for  
part (b) - more direct.

EC (5)

## ADDITIONAL KINETICS Q's

Q Alcohol (ethanol) is converted to acetaldehyde by the enzyme liver alcohol dehydrogenase (LADH)

Clearance of alcohol from the body after consumption is a zero order process, rate constant  $k = 0.20 \text{ oz} \cdot \text{h}^{-1}$

a) If five 10 oz beers are consumed, how long before there is no alcohol left in the body?

Data: 5% alcohol in the beer.

$$50 \text{ oz} \times 0.05 = 2.5 \text{ oz alcohol}$$

$$[\text{Alc.}] = [\text{Alc.}]_0 - 0.20 \text{ oz} \cdot \text{h}^{-1} \times t$$

$$0 = 2.5 - 0.20t \text{ gives } t = \underline{\underline{13 \text{ hours}}}$$

b) If the legal limit of blood alcohol was equivalent to 1 beer, how long after consuming 5 beers would it be safe to drive?

$$1 \text{ beer, } 10 \text{ oz} \times 0.05 = 0.50 \text{ oz. alcohol}$$

$$[\text{Alc.}] = [\text{Alc.}]_0 - 0.20t$$

$$0.50 = 2.5 - 0.20t$$

$$t = \underline{\underline{10 \text{ hours}}}$$

Drinking & driving do not mix!

## ADDITIONAL KINETICS Q's Kin ①

Q. The 100 teeny weeny bacteria sitting on your donut multiply according to:

<u>t/min</u>	<u># of bacteria</u>
0	100
15	200
30	400
45	800
60	1600

a) What is the order of the rate of production of the bacteria?

From data # of bacteria double every 15 min. Doubling is independent of the no. of bacteria.

$\therefore$  rate law is 1st order.

Could confirm by plotting  $\ln(\# \text{ b's})$  against  $t$ .

Half life must be 15 min.

Kin (2)

b) How many bacteria will there be after 2 hours?

- could extrapolate the data.

or 2 hours = 8 half-lives

and every  $\frac{1}{2}$ -life the no. doubles.

$$\therefore \text{we have } \# = 100(2)(2)(2)(2)(2)(2)(2)(2)$$

$$\begin{aligned} \text{OR do part c) first} &= 100(2)^8 \\ \text{and use} &= 2.56 \times 10^4 \text{ (yummy!)} \\ \ln[A] &= \ln[A]_0 - kt \end{aligned}$$

Note: concn A  $\gg$  [A]<sub>0</sub>, watch signs

c) What is the rate constant for the process?

1st order rx.

- reciprocal relationship between  $t_{1/2}$  & k

$$t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$$

$$k = \frac{0.693}{t_{1/2}} = \frac{0.693}{15 \text{ min}} = 4.6 \times 10^{-2} \text{ min}^{-1}$$

Kin (3)

Q

A kinetic investigation of the reaction



was carried out. Beginning with equal concentrations of A and B, the rate after 50% of A had reacted was 37.5% of the initial rate.

Determine rate law

---

<u>% consumption of A</u>	<u>[A]</u>	<u>[B]</u>	<u>Rate.</u>
0%	$[A]_0 = x$	$[B]_0 = x$	R.
50%	$(1 - 0.5)x$	$(1 - \frac{1}{2}(0.5))x$	0.375 R
	↓ $0.5x$	↓ $0.75x$	0.375 R

Kin (4)

$$\text{Rate} = k[A]^m[B]^n$$

$$R = k[x]^{m+n} \quad \text{and}$$

$$0.375R = k[0.5x]^m[0.75x]^n$$

Let initial rate = 100, and initial concn of  $x = 1$

$$100 = k[1]^{m+n} \quad \therefore k = 100.$$

$$37.5 = 100[0.5]^m[0.75]^n.$$

$$0.375 = [0.5]^m[0.75]^n.$$

$m = n = 1$  only values that will satisfy this

---

Q. Rx. of A  $\rightarrow$  products

$$t_{1/2} = 12.0 \text{ min}$$

$$[A]_0 = 0.100 \text{ mol L}^{-1}$$

What is the time for  $[A] = 0.020 \text{ mol L}^{-1}$

for a) 1st order rx b) 2nd order rx?

a) First order

$$t_{1/2} = \frac{0.693}{k}$$

$$\therefore k = \frac{0.693}{12 \text{ min}} = 5.78 \times 10^{-2} \text{ min}^{-1}$$

$$\ln[A] = \ln[A]_0 - kt ; \ln\left(\frac{[A]}{[A]_0}\right) = -kt$$

$$\ln\left(\frac{0.020}{0.100}\right) = -5.78 \times 10^{-2} \times t$$

$$t = \underline{\underline{27.8 \text{ min}}}$$

b) Second order

$$t_{1/2} = \frac{1}{k[A]_0}$$

$$\therefore 12 = \frac{1}{k(0.100)} ; k = 0.833 \text{ M}^{-1} \text{ s}^{-1}$$

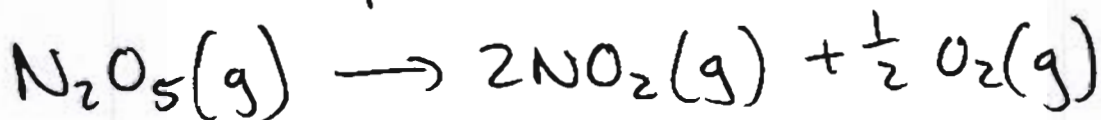
$$\frac{1}{[A]} = \frac{1}{[A]_0} + kt$$

$$\therefore \frac{1}{0.020} - \frac{1}{0.100} = 0.833 \times t$$

$$t = \underline{\underline{48 \text{ min}}}$$

Kin ⑥

Q The decomposition of  $N_2O_5(g)$ :



has the rate law

$$-\frac{d[N_2O_5]}{dt} = k[N_2O_5] \text{ with } k = 1.5 \times 10^{-3} \text{ s}^{-1} \text{ at } 328 \text{ K}$$

↑  
rate

a) Calc the half-life

$$\ln 2 = kt_{\frac{1}{2}} ; \quad t_{\frac{1}{2}} = \frac{0.693}{k}$$
$$= \frac{0.693}{1.5 \times 10^{-3} \text{ s}^{-1}} = \underline{\underline{4.62 \times 10^2 \text{ s}}}$$

b) If the initial concn of  $N_2O_5$  is  $1.46 \times 10^{-2} \text{ mol dm}^{-3}$ , calc the concns of  $N_2O_5$ ,  $NO_2$  and  $O_2$  after 10 min.

$$\ln \frac{[N_2O_5]_0}{[N_2O_5]} = kt = (1.5 \times 10^{-3} \text{ s}^{-1})(10 \times 60 \text{ s})$$
$$= 9.00 \times 10^{-1} \quad \text{Kin. (7)}$$

$$\therefore \frac{[N_2O_5]_0}{[N_2O_5]} = 2.46$$

$$\therefore [N_2O_5] = \frac{1.46 \times 10^{-2} \text{ mol dm}^{-3}}{2.46} = \underline{\underline{5.94 \times 10^{-3} \text{ mol dm}^{-3}}}$$

Change in concn of  $N_2O_5$  is  $\therefore$ ,  
 $(14.6 \times 10^{-3}) - (5.94 \times 10^{-3}) = 8.66 \times 10^{-3} \text{ mol dm}^{-3}$

i.e. have lost this much  $N_2O_5$

$$\begin{aligned} \therefore [NO_2] &= 2 \times 8.66 \times 10^{-3} \\ &= \underline{\underline{1.7 \times 10^{-2} \text{ mol dm}^{-3}}} \end{aligned}$$

$$\begin{aligned} [O_2] &= \frac{1}{2} (8.66 \times 10^{-3}) \\ &= \underline{\underline{4.3 \times 10^{-3} \text{ mol dm}^{-3}}} \end{aligned}$$

c) Calculate the time required for 80% reaction of  $N_2O_5$

$$\text{At 80\% rx.}, \quad \frac{[N_2O_5]_0}{[N_2O_5]} = \frac{100}{20} = 5$$

$$\ln \frac{[N_2O_5]_0}{[N_2O_5]} = kt; \quad \ln 5 = 1.50 \times 10^{-3} \text{ s}^{-1} \times t$$

$$t = \underline{\underline{1.07 \times 10^3 \text{ s}}} \quad (17.9 \text{ min})$$

d) Starting with  $1.46 \times 10^{-2} \text{ mol dm}^{-3}$   $N_2O_5$ , calculate the time required for the rate to drop to  $1.00 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}$

$$\text{Rate} = k[N_2O_5]$$

$$[N_2O_5] = \frac{1.00 \times 10^{-5} \text{ mol dm}^{-3} \text{ s}^{-1}}{1.50 \times 10^{-3} \text{ s}^{-1}} = 6.67 \times 10^{-3} \text{ mol dm}^{-3}$$

Kin (9)

$$\ln \frac{[\text{N}_2\text{O}_5]_0}{[\text{N}_2\text{O}_5]} = kt \quad ; \quad t = \frac{1}{k} \ln \frac{[\text{N}_2\text{O}_5]_0}{[\text{N}_2\text{O}_5]}$$

$$t = \frac{1}{1.50 \times 10^{-3} \text{ s}^{-1}} \ln \frac{1.46 \times 10^{-2}}{6.67 \times 10^{-3}}$$

$$t = \underline{\underline{5.23 \times 10^2 \text{ s}}}$$

e) Starting with  $1.46 \times 10^{-2} \text{ mol dm}^{-3} \text{ N}_2\text{O}_5$   
calc. the rate after 20.0 min

$$\ln \frac{[\text{N}_2\text{O}_5]_0}{[\text{N}_2\text{O}_5]} = kt = (1.5 \times 10^{-3} \text{ s}^{-1})(20.0 \times 60 \text{ s})$$
$$= 1.80$$

$$\frac{[\text{N}_2\text{O}_5]_0}{[\text{N}_2\text{O}_5]} = e^{1.80} = 6.05$$

$$[\text{N}_2\text{O}_5] = \frac{1.46 \times 10^{-2} \text{ mol dm}^{-3}}{6.05}$$

$$\text{Rate} = k[\text{N}_2\text{O}_5] = (1.50 \times 10^{-3} \text{ s}^{-1}) \times \frac{1.46 \times 10^{-2}}{6.05}$$
$$= \underline{\underline{3.62 \times 10^{-6} \text{ mol dm}^{-3} \text{ s}^{-1}}}$$

Km(10)