Big Idea 3: Changes in matter involve the rearrangement and/or reorganization of atoms and/or the transfer of electrons

Chemical reactions ☺

Enduring Understandings

- 3.A: Chemical changes are represented by a balanced chemical equation that identifies the ratios with which reactants react and products form.
- 3.B: Chemical reactions can be classified by considering what the reactants are, what the products are, or how they change from one into the other. Classes of chemical reactions include synthesis, decomposition, acid-base, and oxidation-reduction reactions.
- 3.C: Chemical and physical transformations may be observed in several ways and typically involve a change in energy.

Types of Reactions

- Synthesis: $X + Y \Rightarrow XY$
- Decomposition: $XY \Rightarrow X + Y$
- Solubility Reactions: $X^+ + Y^- \Rightarrow XY$
- Acid Base:
  - Hydrolysis: $HX + H_2O \Rightarrow H_3O^+ + X^-$
  - Neutralization: $HX + MOH \Rightarrow MX + HOH$
- Oxidation Reduction
  - Metals oxidize to form cations. $Cu \Rightarrow Cu^{2+} + 2e^-$
  - Non-metals reduce to form anions. $F_2 + 2e^- \Rightarrow 2F^-$
    - $Cu + F_2 \Rightarrow CuF_2$

Solubility

- Different combinations of different ions have different solubility's.

Mixing two solutions of .5M NaCl and .5M $Pb(NO_3)_2$ will produce an ionic solid.

- What is the identity of the solid?
- Write out the molecular equation.
- Write out the ionic equation.
- Write out the net-ionic equation.

What is the identity of the solid?

- Technically this is not required by the AP curriculum… but you should still be able to determine which ionic compound will precipitate.

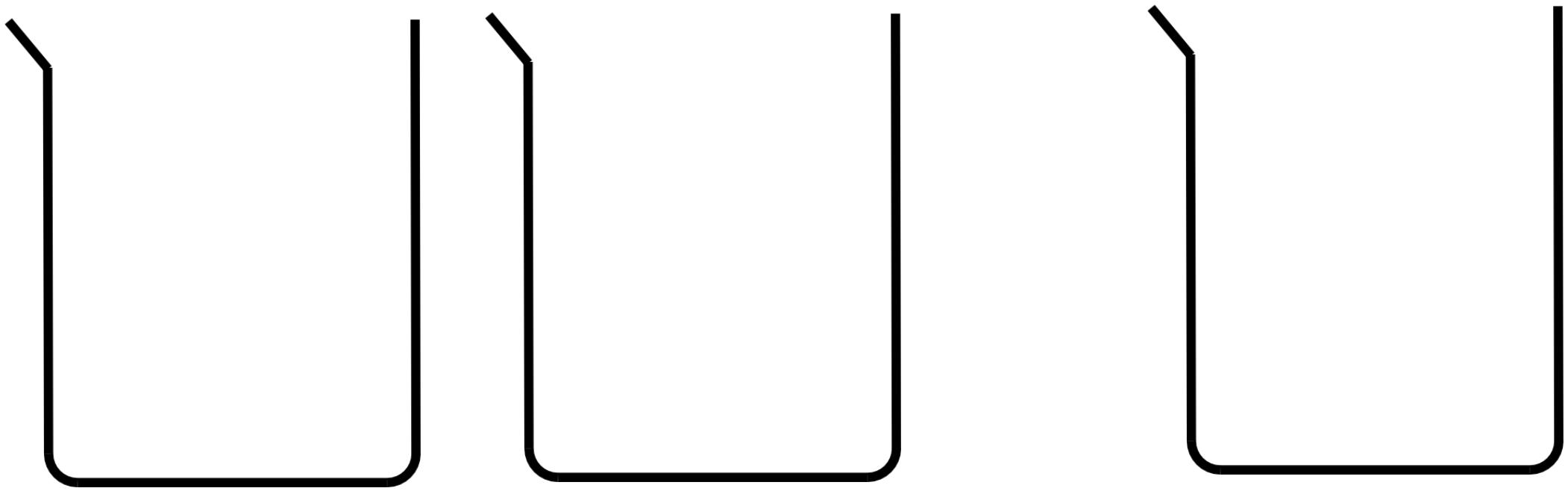
- Due to the high solubility of Na^+ and NO_3^- ions… $PbCl_2$ will be the solid that precipitates.

Writing out chemical reactions!

- Molecular Equation
  - $Pb(NO_3)_2 + NaCl \Rightarrow NaNO_3 + PbCl_{2(s)}$
- Ionic equation
  - $Pb^{+2} \quad 2NO_3^{-1} + 2Na^{+} + 2Cl^{-1} \Rightarrow 2Na^{+} + 2NO_3^{-1} + PbCl_{2(s)}$
- Net ionic equation (removes spectator ions)
  - $Pb^{+2} + 2Cl^{-1} \Rightarrow PbCl_{2(s)}$

Can you draw a picture of this reaction?

- $Pb(NO_3)_2 + NaCl \Rightarrow NaNO_3 + PbCl_{2(s)}$



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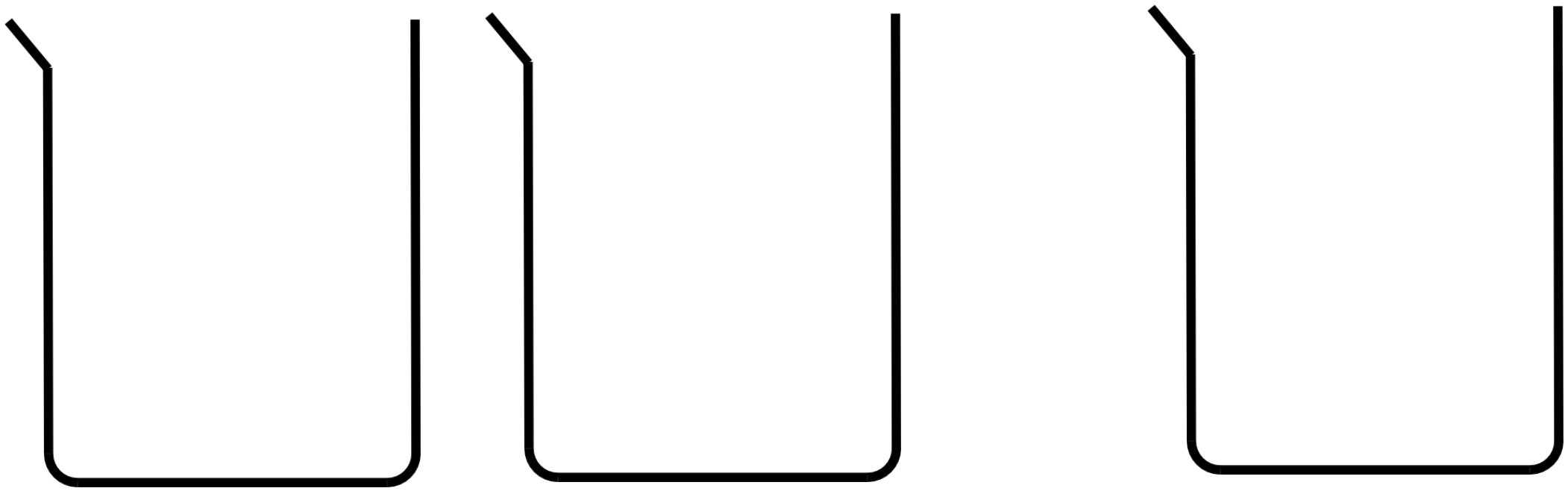


Acid Base

- Hydrolysis: Might be the most important Acid base reaction.
- $HX + H_2O \Rightarrow H_3O^+ + X^-$
- $X^- + H_2O \Rightarrow OH^- + HX$

This is the reactions that produce H^+ and OH^- !

Remember: Ka and Kb are based upon these reactions.

Strong acid vs. weak acids

- Strong acids go to completion
- HCl, HBr, HI, $HClO_4$, HNO_3 H_2SO_4

- What is the pH of a .10M HCl solution?

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

I	.10	-	0	0
Δ	-.10	-	+.10	+.10
E	0	-	+.10	+.10

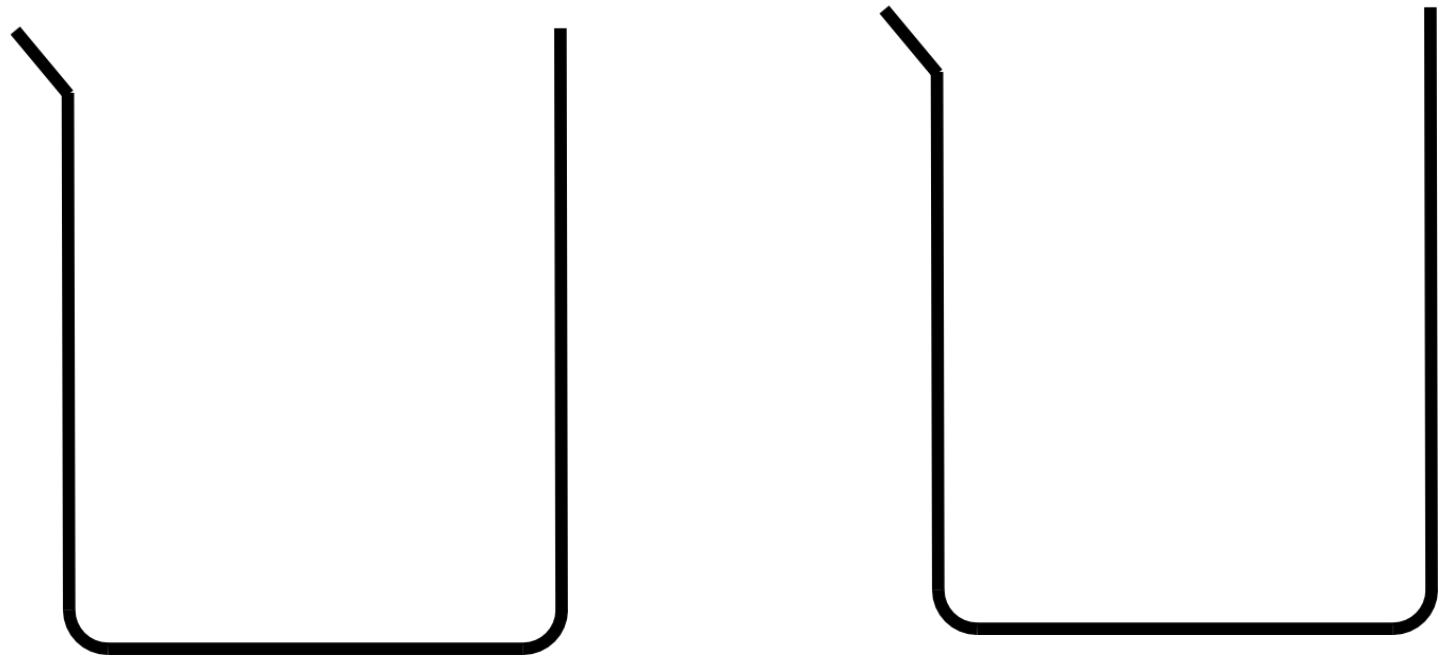
-log [.10] = pH = 1

Don't forget $[H^+]$ = pH

But $[OH^-]$ = pOH

Can you draw a picture of this reaction?

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$



Can you determine the pH of a weak acid?

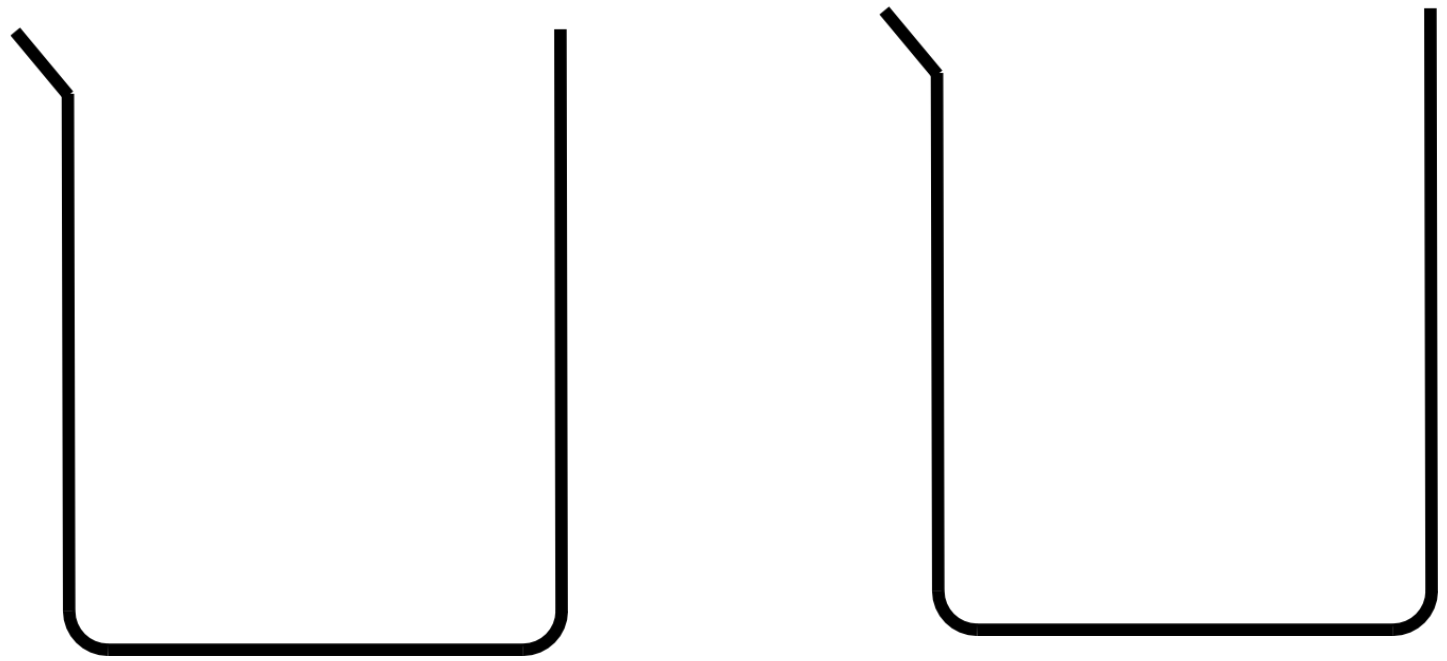
- .10M HX Remember you will need a Ka for this! (Ka = 1.5E-5)

$$HX + H_2O \rightarrow \ H_3O^+ \ + \ X^-$$

Can you determine the percent ionization of this weak acid?

Can you draw a picture?

$$HX + H_2O \rightarrow H_3O^+ + X^-$$



Neutralization reactions

$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + H_2O$$

Example:

$$HF + NaOH \Rightarrow H_2O + NaF$$

Neutralization reactions are used extensively in titrations.

Q: What is the pH of the equivalence point? In other words, what is the acidity of NaF?

Oxidation-reduction

- 1st: You must be able to recongize if a reaction is in fact a redox reaction.

- Assign oxidation states, check reactants and products to see if charges change.
- If they change you should be able to determine who is being oxidized and who is being reduced!

Example: Redox reaction

- Incomplete combustion is a common means to produce carbon monoxide. Carbon monoxide could be burned again to form carbon dioxide.
- Write out the reaction.

- Determine oxidation states.

- Who is being reduced?

Example: Redox reaction

- Incomplete combustion is a common means to produce carbon monoxide. Carbon monoxide could be burned again to form carbon dioxide.
- Write out the reaction.

$CO + O_2 \Rightarrow CO_2$

- Determine oxidation states.

+2-2 0 +4/-2

- Who is being reduced?
- Who gained electrons/reduced: O: 0 to -2
- Who lost electrons/oxidized: C: +2 to +4

Simple redox

- My definition for a typical redox reaction.
- Copper metal + Silver nitrate
  - Write out this reaction
  - Balance using half reactions.
  - Determine voltage
  - Who is being oxidized?
  - Who is being reduced?

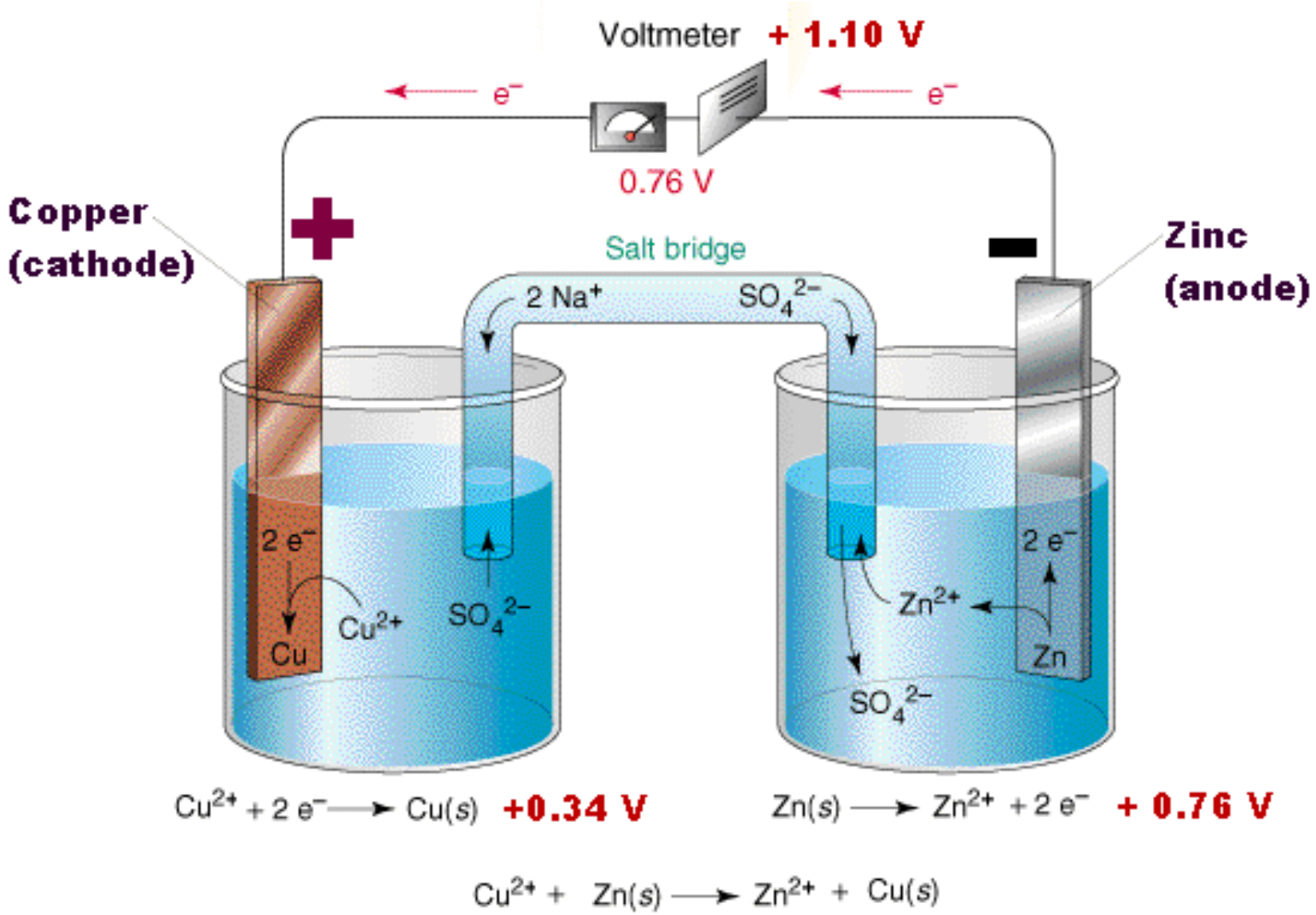
Simple redox

- My definition for a typical redox reaction.
- Copper metal + Silver nitrate
  - Write out this reaction
    - $Cu + 2AgNO_3 \Rightarrow Ag + Cu(NO_3)_2$ Molecular version
    - $Cu + 2Ag^{+} + 2NO_3^{-1} \Rightarrow Ag + Cu^{+2} + 2NO_3^{-1}$ ionic version
    - $Cu + 2Ag^{+} \Rightarrow Ag + Cu^{+2}$
    - Net ionic version
  - Balance using half reactions.
    - $Ag^{+} + 1e \Rightarrow Ag$ +.8V (Reduction) Total voltave = +.46V
    - $Cu \Rightarrow Cu^{+2} + 2e^{-}$ -.34V (Oxidation)
    - remember to flip sign from reduction potential chart

Building an galvanic cell

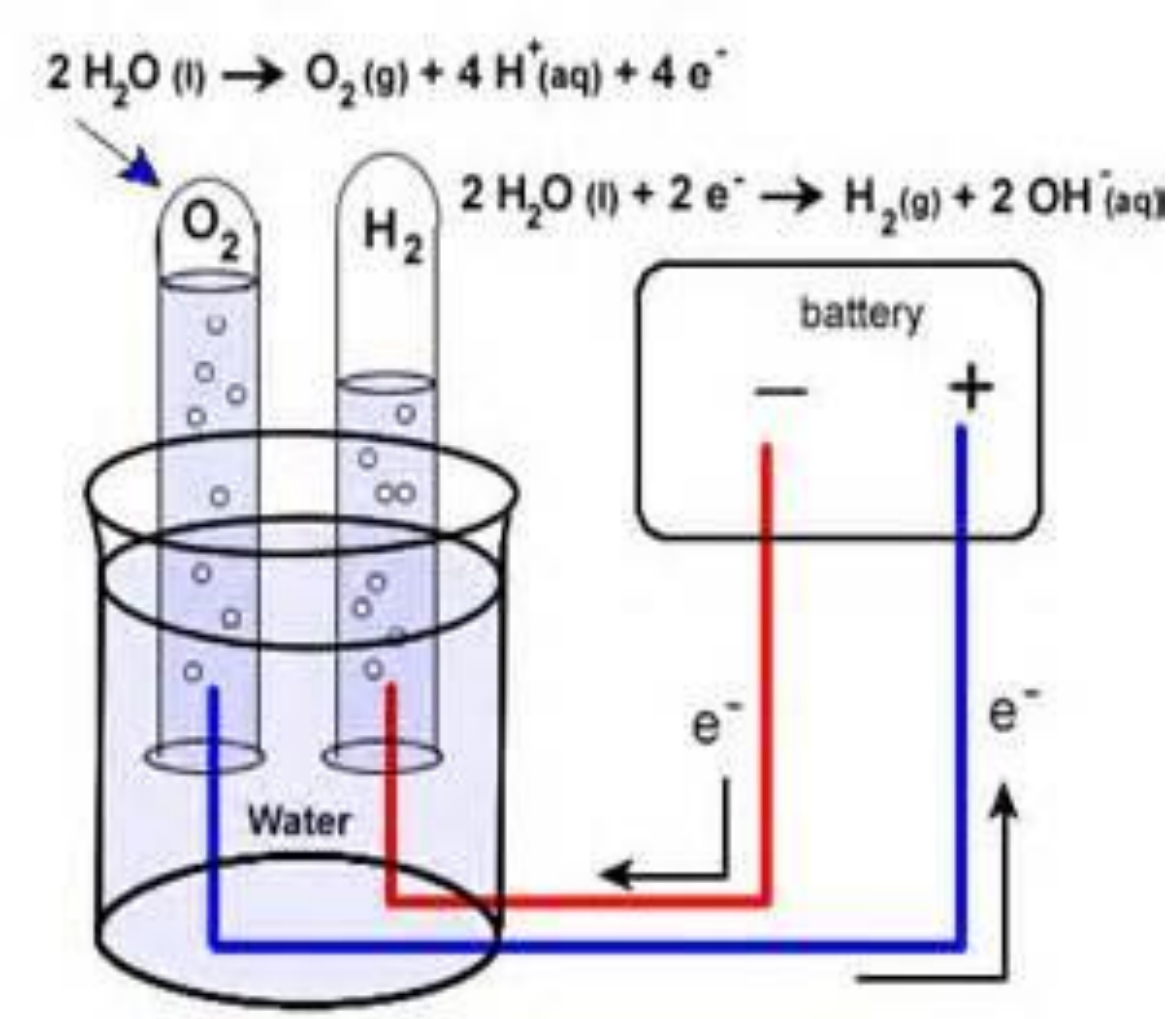
- Which post is gaining mass?
- Which solution is gaining concentration?

What happens if you increase the concentration of each solution?

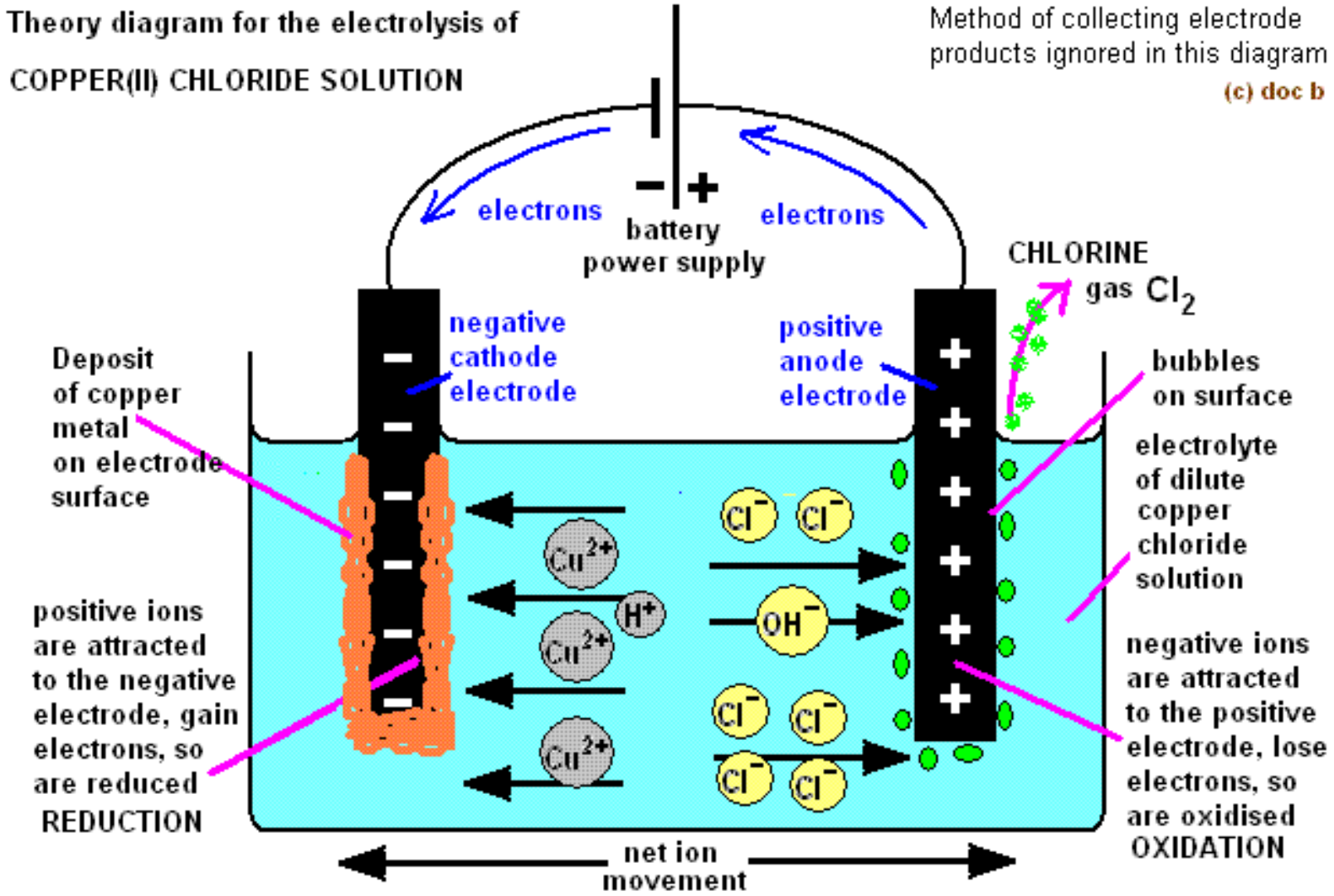


Electrolytic reactions

- I = q/time
- 96500 C = 1 mole of electrons
- Be able to determine produced over time



Alternate example



- If this reaction runs at 2A for 15 min how much volume of Chlorine gas will be produced at STP?

answer

- $2Cl^- \Rightarrow Cl_2 + 2e^-$

- 2.00A for 15.0 min (900.0 sec)
- Determine charge transfered
  - I = q/time I * time = q 2 * 900 sec = 1800 C
- Determine electrons transferred
  - 1800C (1mole e^- /96500 C) = .0186 mole e^-
- Determine moles and volume of chlorine produced.
- .0186 mole e^- (1 mol Cl_2 / 2 mol e) * (22.4L/1mol Cl_2) = .209L

$$H_2C_2O_4(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + HC_2O_4^-(aq)$$

$$HC_2O_4^-(aq) + H_2O(l) \rightleftarrows H_3O^+(aq) + C_2O_4^{2-}(aq)$$

$$H_2O(l) + H_2O(l) \rightleftarrows H_3O^+(aq) + OH^-(aq)$$

11. All the reactions represented above occur in an aqueous solution of oxalic acid. Which of the following represent a Brønsted-Lowry conjugate acid-base pair?

(A) $H_2C_2O_4(aq)$ and $C_2O_4^{2-}(aq)$

(B) $HC_2O_4^-(aq)$ and $C_2O_4^{2-}(aq)$

(C) $HC_2O_4^-(aq)$ and $H_2O(aq)$

(D) $H_3O^+(aq)$ and $OH^-(aq)$

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(C) $HC_2O_4^-(aq)$ and $H_2O(aq)$

(D) $H_3O^+(aq)$ and $OH^-(aq)$

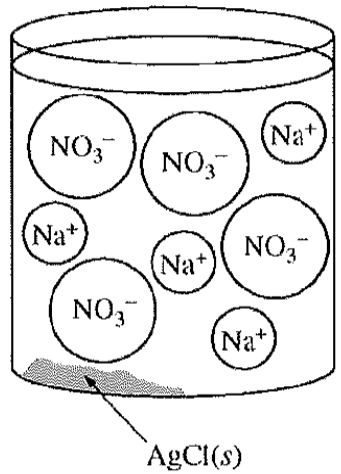
- A few things to keep in mind:
- $H_2C_2O_4$ is diprotic
- Each reaction has an acid base conjugate pair.
- Must be in the same reaction
- B = answer

Q: 15 covered in acid base equlibrium

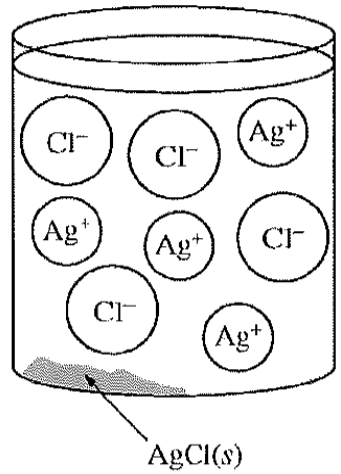
$$Ag^+(aq) + Cl^-(aq) \rightleftarrows AgCl(s)$$

28. A student mixes dilute $AgNO_3(aq)$ with excess $NaCl(aq)$ to form $AgCl(s)$, as represented by the net ionic equation above. Which of the diagrams below best represents the ions that are present in significant concentrations in the solution? (K_{sp} for AgCl is 1.8×10^{-10}.)

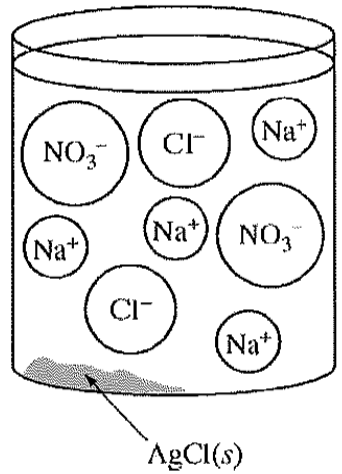
(A)



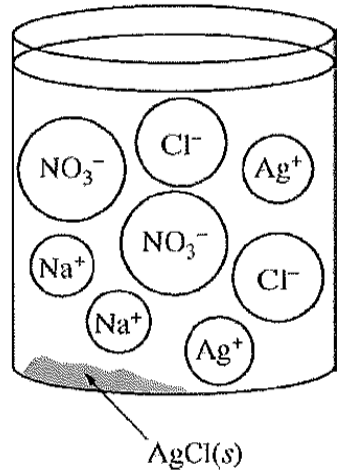
(B)



(C)



(D)



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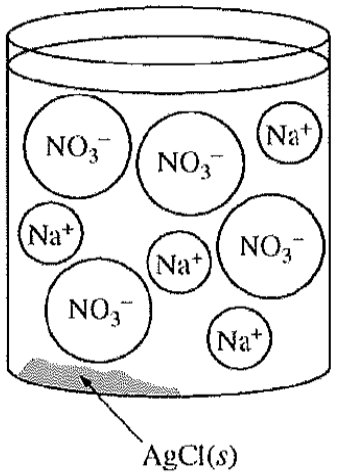

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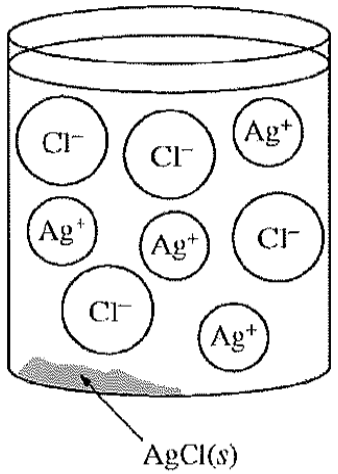
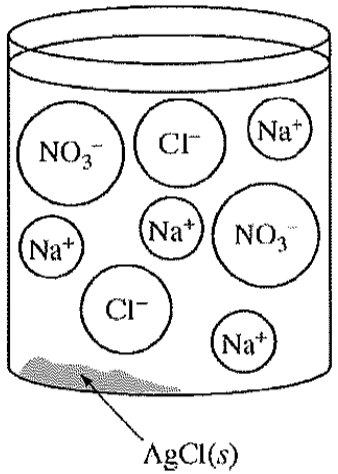
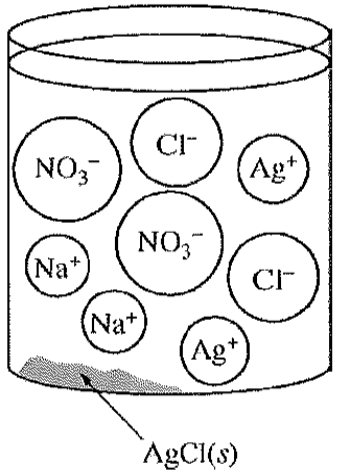


(C)



(D)



- Note the word excess in the problem NaCl… So all Ag^+ is gone.
- B & D is eliminated
- Excess NaCl means there will be some Cl^- in the solution as well.
- A is eliminated.
- C is correct

a #3: Reactions

$$PCl_5(g) \rightleftarrows PCl_3(g) + Cl_2(g)$$

$PCl_5(g)$ decomposes into $PCl_3(g)$ and $Cl_2(g)$ according to the equation above. A pure sample of $PCl_5(g)$ is placed in a rigid, evacuated 1.00 L container. The initial pressure of the $PCl_5(g)$ is 1.00 atm. The temperature is held constant until the $PCl_5(g)$ reaches equilibrium with its decomposition products. The figures below show the initial and equilibrium conditions of the system.

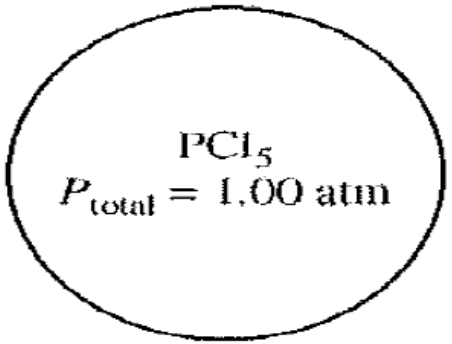


Figure 1: Initial

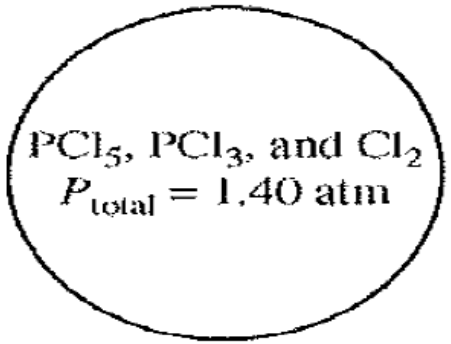


Figure 2: Equilibrium

31. If the decomposition reaction were to go to completion, the total pressure in the container would be

(A) 1.4 atm
(B) 2.0 atm
(C) 2.8 atm
(D) 3.0 atm

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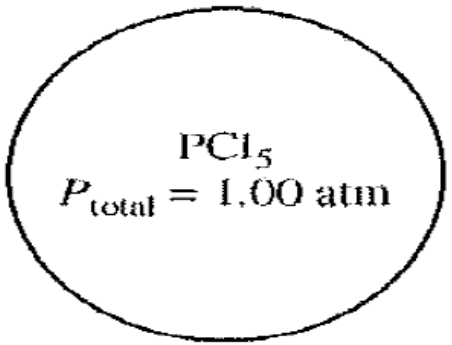

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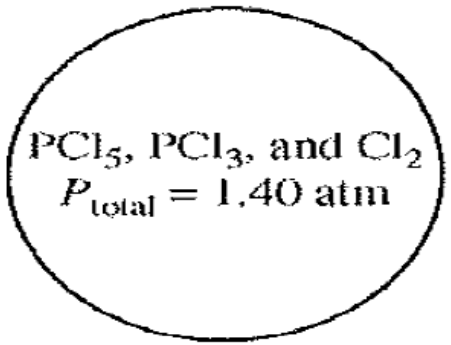

Figure 2: Equilibrium

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(A) 1.4 atm
(B) 2.0 atm
(C) 2.8 atm
(D) 3.0 atm

Going to completions… 1 atm doubles to 2 atm
B is correct

$$K(s) + \frac{1}{2}Cl_2(g) \rightarrow KCl(s) \qquad \Delta H^\circ = -437\ \text{kJ/mol}_{rxn}$$

The elements K and Cl react directly to form the compound KCl according to the equation above. Refer to the information above and the table below to answer the questions that follow.

Process	ΔH° (kJ/mol_{rxn})
$K(s) \rightarrow K(g)$	v
$K(g) \rightarrow K^+(g) + e^-$	w
$Cl_2(g) \rightarrow 2\,Cl(g)$	x
$Cl(g) + e^- \rightarrow Cl^-(g)$	y
$K^+(g) + Cl^-(g) \rightarrow KCl(s)$	z

35. What remains in the reaction vessel after equal masses of $K(s)$ and $Cl_2(g)$ have reacted until either one or both of the reactants have been completely consumed?

(A) KCl only
(B) KCl and K only
(C) KCl and Cl_2 only
(D) KCl, K, and Cl_2

$K(s) + 1/2Cl_2 \Rightarrow KCl$

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(A) KCl only
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- This is a stoichiometric proportion problem.
- Equal masses. In my mind I choose 40 grams.
- 40g = 1 mole of K (40g/mol)
- 40 g = (70g/mol) > .57 moles
  - Little more then a half a mole.

Since K is used up twice as fast it will run out 1st

Excess = Cl_2 and product KCl.

Answer is C

Half-cell 1: strip of $Al(s)$ in 1.00 M $Al(NO_3)_3(aq)$

Half-cell 2: strip of $Cu(s)$ in 1.00 M $Cu(NO_3)_2(aq)$

Half-cell 3: strip of $Fe(s)$ in 1.00 M $Fe(NO_3)_2(aq)$

Galvanic Cell	Half-cells	Cell Reaction	E°_{cell} (V)
X	1 and 2	$2\,Al(s) + 3\,Cu^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Cu(s)$	2.00
Y	1 and 3	$2\,Al(s) + 3\,Fe^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Fe(s)$	1.22
Z	2 and 3	$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$	?

45. What is the standard cell potential of galvanic cell Z?

(A) 0.26 V
(B) 0.78 V
(C) 2.34 V
(D) 3.22 V

46. In galvanic cells Y and Z, which of the following takes place in half-cell 3 ?

(A) Reduction occurs in both cell Y and cell Z.
(B) Oxidation occurs in both cell Y and cell Z.
(C) Reduction occurs in cell Y, and oxidation occurs in cell Z.
(D) Oxidation occurs in cell Y, and reduction occurs in cell Z.

47. If the half-cell containing 1.00 M $Fe(NO_3)_2(aq)$ in galvanic cells Y and Z is replaced with a half-cell containing 5.00 M $Fe(NO_3)_2(aq)$, what will be the effect on the cell voltage of the two galvanic cells?

(A) The voltage will increase in both cells.
(B) The voltage will decrease in both cells.
(C) The voltage will increase in cell Y and decrease in cell Z.
(D) The voltage will decrease in cell Y and increase in cell Z.

Half-cell 1: strip of $Al(s)$ in 1.00 M $Al(NO_3)_3(aq)$
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Adjusting cell X and Z will be needed to add up to Z. Then add up the Ecells to get Z

- X cell can stay as is +2.00V
- Y cell needs to get flipped -1.22 V New voltage = +.78V B is correct answer

This will add up to correct reaction. Stoich ratios do not affect voltage.

Half-cell 1: strip of $Al(s)$ in 1.00 M $Al(NO_3)_3(aq)$

Half-cell 2: strip of $Cu(s)$ in 1.00 M $Cu(NO_3)_2(aq)$

Half-cell 3: strip of $Fe(s)$ in 1.00 M $Fe(NO_3)_2(aq)$

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Iron is half cell 3 (see top)

Y: Fe^{2+} is being reduced to Fe

Z: Fe is being oxidized to Fe^{2+}

C is correct answer

Half-cell 1: strip of $Al(s)$ in 1.00 M $Al(NO_3)_3(aq)$

Half-cell 2: strip of $Cu(s)$ in 1.00 M $Cu(NO_3)_2(aq)$

Half-cell 3: strip of $Fe(s)$ in 1.00 M $Fe(NO_3)_2(aq)$

Galvanic Cell	Half-cells	Cell Reaction	E°_{cell} (V)
X	1 and 2	$2\,Al(s) + 3\,Cu^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Cu(s)$	2.00
Y	1 and 3	$2\,Al(s) + 3\,Fe^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Fe(s)$	1.22
Z	2 and 3	$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$	?

47. If the half-cell containing 1.00 M $Fe(NO_3)_2(aq)$ in galvanic cells Y and Z is replaced with a half-cell containing 5.00 M $Fe(NO_3)_2(aq)$, what will be the effect on the cell voltage of the two galvanic cells?

(A) The voltage will increase in both cells.
(B) The voltage will decrease in both cells.
(C) The voltage will increase in cell Y and decrease in cell Z.
(D) The voltage will decrease in cell Y and increase in cell Z.

Half-cell 1: strip of $Al(s)$ in 1.00 M $Al(NO_3)_3(aq)$
Half-cell 2: strip of $Cu(s)$ in 1.00 M $Cu(NO_3)_2(aq)$
Half-cell 3: strip of $Fe(s)$ in 1.00 M $Fe(NO_3)_2(aq)$

Galvanic Cell	Half-cells	Cell Reaction	E°_{cell} (V)
X	1 and 2	$2\,Al(s) + 3\,Cu^{2+}(aq) \rightarrow$ [illegible] $+ 3\,Cu(s)$	2.00
Y	1 and 3	$2\,Al(s) + 3\,Fe^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Fe(s)$	1.22
Z	2 and 3	$Fe(s) + Cu^{2+}(aq) \rightarrow Fe^{2+}(aq) + Cu(s)$	?

47. If the half-cell containing 1.00 M $Fe(NO_3)_2$([illegible]) in galvanic cells Y and Z is replaced with a half-cell containing 5.00 M $Fe(NO_3)_2(aq)$, what will be the effect on the cell voltage of the two galvanic cells?

(A) The voltage will increase in both cells.
(B) The voltage will decrease in both cells.
(C) The voltage will increase in cell Y and decrease in cell Z.
(D) The voltage will decrease in cell Y and increase in cell Z.

Note: In Y Fe^{2+} is a reactant whereas in Z it is a product.

Increasing a reactions will drive up Voltage
Increasing a product will drive down Voltage.

Y will increase
Z will decrease

C = answer