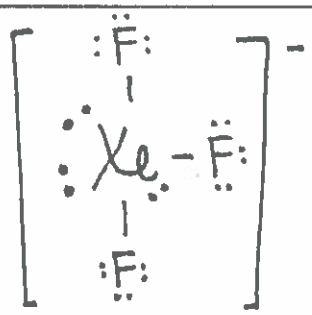
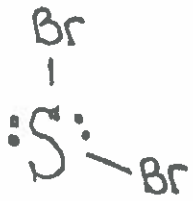
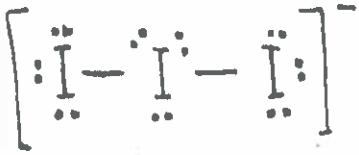
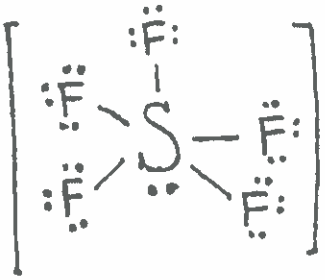
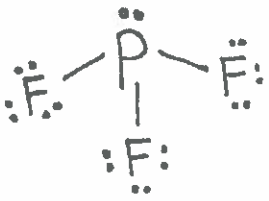
WEEK 2 KEY

Compound	Lewis Structure	Electronic Geometry	Molecular Geometry, Bond Angle	Is this molecule polar?	Hybridization of central atom
XeF_3^-		octahedral	T-shaped 90°	yes	sp^3d^2
SBr_2		tetrahedral	bent <109.5°	yes	sp^3
I_3^-		trigonal bipyramidal	linear 180°	no	sp^3d
SF_5		octahedral	square pyramidal 90°	yes	sp^3d^2
PF_3		tetrahedral	trigonal pyramidal <109.5°	yes	sp^3

1. The data to the right was collected for the reaction below. What is the <u>average rate of ammonia (NH_3) formation</u> from 30-60s if this experiment was performed in a 2.00 L flask?

$$2N_2\,(g) + 3H_2\,(g) \rightarrow 2NH_3\,(g)$$

Time (s)	Moles of H_2
0	8.54
15	7.44
30	6.29
45	5.16
60	4.09

$$\frac{\Delta[H_2]}{\Delta t} = \frac{4.09\,mol/2L - 6.29\,mol/2L}{60s - 30s} = -0.0366667\ Ms^{-1}$$

$$rate = \frac{-1}{3}\frac{\Delta[H_2]}{\Delta t} = \frac{1}{2}\frac{\Delta[NH_3]}{\Delta t}$$

$$\frac{\Delta[NH_3]}{\Delta t} = \frac{-2}{3}\cdot\frac{\Delta[H_2]}{\Delta t} = \frac{-2}{3}(-.0366667) = 2.44\times10^{-2}\ Ms^{-1}$$

2. The rate of the reaction, $HgCl_2(aq) + \frac{1}{2}C_2O_4^{2-}(aq) \rightarrow Cl^-(aq) + CO_2(g) + \frac{1}{2}Hg_2Cl_2(s)$, is followed by measuring the number of moles of Hg_2Cl_2 that precipitate per liter per second. What is the rate constant and overall rate order?

$[HgCl_2]$	$[C_2O_4^{2-}]$	Initial Rate (mol/L ·s)
0.10	0.10	1.3×10^{-7}
0.10	0.20	5.2×10^{-7}
0.20	0.20	1.0×10^{-6}
0.20	0.10	2.6×10^{-7}

$\frac{exp\,2}{exp\,1}$: $\frac{5.2\times10^{-7}Ms^{-1}}{1.3\times10^{-7}Ms^{-1}} = \frac{k(0.10)^m(0.20)^n}{k(0.10)^m(0.10)^n}$

$4 = 2^n$

$\therefore n = 2$, $C_2O_4^{2-}$ is 2nd order

$\frac{exp\,4}{exp\,1}$: $\frac{2.6\times10^{-7}Ms^{-1}}{1.3\times10^{-7}Ms^{-1}} = \frac{k(0.20)^m(0.10)^n}{k(0.10)^m(0.10)^n}$

$2 = 2^m$

$\therefore m = 1$, $HgCl_2$ is 1st order

overall rate order $= m + n = 1 + 2$

overall rate order = 3

$rate = k[HgCl_2][C_2O_4^{2-}]^2$

$1.3\times10^{-7}\ Ms^{-1} = k(0.10M)(0.10M)^2$

$\therefore k = 1.3\times10^{-4}\ M^{-2}s^{-1}$

3. The rate law for the reaction: $NH_4^+(aq) + NO_2^-(aq) \rightarrow N_2(g) + 2H_2O(l)$ is given by rate = $k[NH_4^+][NO_2^-]$. At 25°C, the rate constant is $3.0\times10^{-4}\ M^{-1}\ s^{-1}$. Calculate the rate of reaction at this temperature if $[NH_4^+] = 0.36$ M and $[NO_2^-] = 0.075$ M. What is the rate of H_2O production?

$$\text{rate} = k[NH_4^+][NO_2^-]$$

$$= 3.0\times10^{-4}\ M^{-1}S^{-1}(0.36M)(0.075M)$$

$$\text{rate} = 8.1\times10^{-6}\ MS^{-1}$$

$$\text{rate} = \Delta[H_2O]/2\Delta t$$

$$\frac{\Delta[H_2O]}{\Delta t} = 1.6\times10^{-5}\ MS^{-1}$$

4. Determine the overall orders of the reactions to which the following rate laws apply:

a. rate = $k[NO_2]^2$

second order

b. rate = k

zero order

c. rate = $k[H_2]^2[Br_2]^{1/2}$

2.5 order

d. rate = $k[NO]^2[O_2]$

third order

5. Write the reaction rate expressions for the following reactions in terms of the disappearance of the reactants and the appearance of products:

a. $2NO_2 \rightarrow 2NO + O_2$

$$\text{rate} = \frac{-\Delta[NO_2]}{2\Delta t} = \frac{\Delta[NO]}{2\Delta t} = \frac{\Delta[O_2]}{\Delta t}$$

b. $4NH_3 + 5O_2 \rightarrow 4NO + 6H_2O$

$$\text{rate} = \frac{-\Delta[NH_3]}{4\Delta t} = \frac{-\Delta[O_2]}{5\Delta t} = \frac{\Delta[NO]}{4\Delta t} = \frac{\Delta[H_2O]}{6\Delta t}$$

c. $2NO_2(g) + Cl_2(g) \rightarrow 2NO_2Cl(g)$

$$\text{rate} = \frac{-\Delta[NO_2]}{2\Delta t} = \frac{-\Delta[Cl_2]}{\Delta t} = \frac{\Delta[NO_2Cl]}{2\Delta t}$$