1. a volume of 1.0 L to 10.1 L
Against an external pressure of 0.50 atm
The gas absorbs 250 J of heat from the surroundings
q, w and ΔU (ΔE)=?

q= +250 j
w= - P × ΔV= -(0.50 atm) × (10.1 L – 1.0 L)= -4.55 atm × (101.3 J/1 L.atm)= -460.915 J
ΔE= ΔU= q + w = +250 J + (-460.915 J)= -210.915 J

2. $\Delta H^o_{ol}[HCl_{(g)}] = -92.31$ kJ
$\Delta H^o_{ol}[CH_{4(g)}] = -74.81$ kJ
$CH_4(g) + 4Cl_2(g) \rightarrow CCl_4(g) + 4HCl(s)$ $\Delta H^o = -397{,}3$ kJ
The standard enthalpy of formation of $CCl_4(g)$ at 25 °C and 1 bar.

$\Delta H^o_{rxn} = [\Delta H^o_{ol}(CCl_4(g) + 4\Delta H^o_{ol}(HCl_{(g)})] - [\Delta H^o_{ol}(CH_{4(g)}) + 4\,\Delta H^o_{ol}(Cl_{2(g)})]$
$-397.3\text{ kJ} = [\Delta H^o_{ol}(CCl_4(g) + 4 \times -92.31\text{ kJ}] - [-74.81\text{ kJ} + 4 \times 0\text{ kJ}]$
$\Delta H^o_{ol}(CCl_4(g)) = -102.87$ kJ

3. 50 g of iron that has an initial temperature of 225 °C and 50 g of gold that has an initial temperature of 25 °C
no heat is lost to the surroundings, what will be the temperature when the two metals reach **thermal equilibrium?**
The specific heat capacity of iron: 0.449 J/g°C and gold: 0.128 J/g°C

$q_{Au} + q_{Fe} = 0$
$q_{Au} = -q_{Fe}$
$q_{Au} = m_{Au} \times S_{Au} \times \Delta t = 50g \times 0.128\text{ J/g°C} \times (t_f - 25\text{ °C})$
$q_{Fe} = m_{Fe} \times S_{Fe} \times \Delta t = 50g \times 0.449\text{ J/g°C} \times (t_f - 225\text{ °C})$
$q_{Au} = -q_{Fe}$
$50g \times 0.128\text{ J/g°C} \times (t_f - 25\text{ °C}) = -(50g \times 0.449\text{ J/g°C} \times (t_f - 225\text{ °C}))$
$t_f = 180.55$ °C

4. increasing boiling point: CH_3CH_2OH, $HOCH_2CH_2OH$, CH_3CH_2Cl, and $ClCH_2CH_2OH$

$CH_3CH_2Cl < CH_3CH_2OH < ClCH_2CH_2OH < HOCH_2CH_2OH$

5. increasing **boiling point** and **vapor pressure**: CH_3OH, He, CH_3Cl and N_2

Boiling point: He< N_2< CH_3Cl< CH_3OH
Vapor pressure: He>N_2> CH_3Cl> CH_3OH

6. Europium crystallizes in a body-centered cubic lattice
The density of Eu is 5.26 g/cm^3
Calculate the unit cell edge length in pm. (Eu MW:152 g/mol).

m= (2 Eu atoms/1 unit cell) × (1 mol Eu/6.022 × 10^{23} Eu atoms) × (152.0 g Eu/1mol Eu) = 5.048 × 10^{-22} g Eu/unit cell
V= (5.048 × 10^{-22} g Eu/1 unit cell) × (1 cm^3/5.26g)= 9.60 × 10^{-23} cm^3/unit cell
a= $V^{1/3}$= (9.60 × 10^{-23} cm^3)$^{1/3}$= 4.58 × 10^{-8} cm=458 pm

7. The density of an aqueous solution containing 10.0 percent of C_2H_5OH by mass is 0.984 g/mL.

Molality, molarity, what volume of the solution would contain 0.125 mole of ethanol?
(M_C*=12.01,* M_H*=1.008,* M_O*=16.00)*

Assume 100 g of solution
0.1 × 100 g= 10g ethanol
100g-10g= 90g=0.090 kg water

10 g ethanol × (1mol/46.07g)= 0.217 mol ethanol
m= mol solute/kg solvent= 0.217mol/0.090kg= 2.41m

100g × (1mL/0.984g)= 102 mL= 0.102 L
M=(mol solute/liters of soln)=0.217mol/0.102L= 2.13 M

0.125mol × (1L/2.13mol)= 0.0587L= 58.7mL

8. 0.100 g of lysozyme is dissolved in 150 g of water at 25 °C.

Calculate the vapor-pressure lowering, the depression in freezing point, the elevation in boiling point, and the osmotic pressure of this solution.
Molar mass of lysozyme = 13930 g/mol
Molar mass of water = 18.02 g/mol
density of the solution = 1 g/mL
$\boldsymbol{P^0_{water}}$ = 23.76 mmHg at 25 °C
$\mathbf{K_f}$ = 1.86 ºC/m
$\mathbf{K_b}$ = 0.52 ºC/m
R = 0.0821 L.atm/mol.K

$n_{lysozyme}$=0.10g × (1mol/13930g)= 7.18 × 10^{-6}mol
n_{water}=150g × (1mol/18.02g)= 8.32 mol

Vapor-pressure lowering:
$\Delta P = X_{lysozyme} \times P^{\circ}_{water} = (n_{lysozyme}/(n_{lysozyme} + n_{water})) \times$ (23.76 mmHg)
=(7.18 × 10^{-6}mol/(7.18 × 10^{-6}mol + 8.32 mol)) × (23.76 mmHg)= 2.05 × 10^{-5} mmHg

Freezing point depression:
$\Delta T_f = K_f \times m$= 1.86 ºC/m × (7.18 × 10^{-6}mol/0.150kg)= 8.90 × 10^{-5} ºC

Boling point elevation:
$\Delta T_b = K_b \times m$= 0.52 ºC/m × (7.18 × 10^{-6}mol/0.150kg)= 2.5 × 10^{-5} ºC

Osmotic pressure:
π=M × R × T= (7.18 × 10^{-6}mol/0.150L) × 0.0821 L.atm/mol.K × (298K)
= 1.17 × 10^{-3} atm=0.889 mmHg

9. increasing solubility in water: O_2, LiCl, Br_2, (CH_3OH)
$O_2 < Br_2 <$ LiCl $< CH_3OH$

10. A + B $\rightarrow$ products

determine the order of the reaction and calculate the rate constant

	[A] (*M*)	**[B] (*M*)**	**Rate (M/s)**
1	1.50	1.50	3.20×10^{-1}
2	1.50	2.50	3.20×10^{-1}
3	3.00	1.50	6.40×10^{-1}

Compare with first and second sets: Changing [B] does not affect the rate of the reaction, zero order in B

Compare with first and third sets: doubling [A] doubles the rate of the reaction, first order in A

Rate= k × [A]

3.20×10^{-1} M/s= k × (1.50 M)

k= 0.212 s^{-1}

$Rate_2/Rate_1 = (3.20 \times 10^{-1}/3.20 \times 10^{-1}) = 1 = (k \times (1.50)^x \times (2.50)^y)/ (k \times (1.50)^x \times (1.50)^y)$

y= 0 Zero order for B

$Rate_3/Rate_1 = (6.40 \times 10^{-1}/3.20 \times 10^{-1}) = 2 = (k \times (3.00)^x \times (1.50)^y)/ (k \times (1.50)^x \times (1.50)^y)$

X= 1 First order for A

11. The rate constant for the second-order reaction is 0.80/*M* . s at 10°C.
Starting with a concentration of 0.086 M, calculate the concentration of NOBr after 22 s.
Calculate the half-life when $[NOBr]_0$= 0.072 M.

$$2NOBr(g) \rightarrow 2NO(g) + Br_2(g)$$

$1/[NOBr]_t = kt + 1/[NOBr]_0$

$1/[NOBr]_t$= (0.80/M.s) (22s) + 1/0.086M

$1/[NOBr]_t$= 29 M^{-1}

$[NOBr]_t$= 0.034M

$t_{1/2} = 1/k[A]_0$= 1/(0.80/M.s) (0.072M)

$t_{1/2}$= 17 s

12. The rate constants of some reactions double with every 10-degree rise in temperature. Assume that a reaction takes place at 295 K and 305 K.
What must the activation energy?

$In(k_1/2k_1) = E_a/R \times ((T_1-T_2)/ T_1.T_2)$

-0.693= E_a/8.314 J/K.mol ((295K-305K)/295K.305K)

$E_a = 5.18 \times 10^4$ J/mol= 51.8 kJ/mol